CHEMISTRY FOR SCHOOLS

MILLS

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CHEMISTRY FOR SCHOOLS

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PREFACE

DURING the last decade an animated discussion has been carried on in the scientific literature as to the proper place of the atomic and molecular hypotheses in chemistry; and the conviction is rapidly gaining ground that these hypotheses should be excluded altogether from the earlier stages of instruction and should be introduced only after the pupil has acquired familiarity with the fundamental conceptions and quantitative laws of the science. In this way only is it possible to achieve that distinction between fact and hypothesis so long recognized as necessary to clear thinking.

Some time ago Messrs. Gage and Company asked us to prepare a book on chemistry treated from this point of view. We felt ourselves unable to comply; we were unwilling to be financially interested in a book, the very necessity of which might be, or might seem to be, traceable to ourselves, that is, to new methods of treatment or new branches of chemistry recommended by us for the curriculum.

When, however, Mr. Mills had prepared the manuscript for a text-book and asked us to revise it, then, since we have a great interest in the teaching of chemistry in the schools of Ontario, we could not refuse his request, and have accordingly done our best, in the limited time at our disposal, by revising and rearranging the subject matter, to bring the book into accord with modern views and to reduce the hypothetical element to a minimum.

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UNIVERSITY OF TORONTO,
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CHEMISTRY FOR SCHOOLS

CHAPTER I.—REACTIONS

SUBSTANCES.

OBJECTS about us are distinguishable because of their properties, such as form, size, hardness, color, odor, taste, density, combustibility, solubility, etc. All of these properties, however, are not of equal interest to the chemist, who concerns himself mainly with those which are not affected by changing the size or form of the object. Tables, chairs, and doors, for instance, have each a definite form, by which we recognize and distinguish them; while with respect to their other properties, density, combustibility, etc., they are the same: if broken to pieces the size and form are altered, while the others remain unchanged. These objects are therefore said to consist of the same "substance," namely, wood. Similarly, if a knife blade, or a window pane, be broken, the "substance" of which each is composed—steel or glass—remains unaltered.

Properties of a Substance.

SUPPLIES: Piece of roll sulphur, graduate, dry cell, electric bell, test tube, carbon disulphide, evaporating dish, lamp, silver coin.

Experiments:

1. Examine a piece of roll sulphur and note its color, its lustre; test its hardness by scratching it with a knife.

2. Weigh a piece of roll sulphur, about 20 grams; partly fill a graduate with water and read the volume, drop in the sulphur and take the reading again. Find the weight of 1 cc. of the sulphur or its specific gravity.

3. Connect the wires leading from the poles of a dry cell to an electric bell, so as to form a complete circuit. Disconnect the wire and place in the circuit a piece of dry sulphur.

4. Rub a piece of roll sulphur, the end of a fountain pen, a rubber comb, a piece of thick glass, or sealing wax, with a dry cloth, or on the coat sleeve, and bring it near a small scrap of paper.

5. Draw the blade of a pocket knife across the end of a bar magnet and bring it close to an iron tack or pin. Repeat with a piece of sulphur.

6. Place a small piece of sulphur in a test tube, add water and shake. In another test tube place a similar piece of sulphur, add carbon disulphide* and shake.

7. Put a piece of sulphur in an evaporating dish and heat over the flame of a spirit lamp.

8. Ignite the melted sulphur in the evaporating dish with a lighted splinter.

9. Rub some flowers of sulphur on a silver coin.

The above experiments serve to acquaint us with the color, lustre, hardness, and specific gravity of sulphur; they show also that it is non-magnetic; that it is not a conductor of electricity; that it may be electrified and melted, and that it blackens silver, and dissolves in carbon disulphide. All these are "properties" of sulphur.

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* Carbon disulphide evaporates rapidly, and the vapor ignites at a low temperature, therefore it must not be brought near a flame of any kind.

Necessity for Quantitative Determination.

In most cases the properties, as determined in the above experiments, are not sufficiently definite to identify a substance. For instance, to distinguish between sulphur and butter by the hardness, it is necessary to know *how* hard each of the materials is; and to distinguish between sulphur and wax by their "solubility" in carbon disulphide, it is necessary to know *how* soluble they are, or how much will dissolve in a certain amount of the liquid. Generally the degree to which a property is possessed by a substance may be indicated by a number, and in such cases the property is said to be expressed quantitatively. In the above-mentioned experiments, the only property that was determined quantitatively was the density. By experiments that will not be described here, the color, power to conduct electricity, melting point, etc., may also be determined quantitatively, and it is only when this has been done that a substance can be considered to be completely described.

CHANGES THAT SUBSTANCES UNDERGO.

In the last section a number of the properties of a substance have been illustrated; the present section will deal with the changes which substances may undergo when heated, electrolysed, or brought together with other materials.

We recognize changes going on everywhere about us, as in the crumbling of rocks, in the growth of animal and plant organisms, in the burning of lamps, wood, coal; in the preparation of food, as in baking and boiling; in the decomposition of organic matter, the rusting of iron, the souring of milk, the fermentation of fruit, the process of digestion, etc., and it is the task of chemistry to find out the laws of all such phenomena.

SUPPLIES: Spirit lamp, retort stand, flask, test tubes, rubber cork, glass tubing, platinum, copper and iron wire, sugar, alum, bluestone (copper sulphate), and sawdust.

Experiments :

1. Heat a piece of platinum wire in the flame of a spirit lamp. What properties change in the flame ?

Let the wire cool. Has the platinum undergone any permanent change ?

2. Repeat with copper or iron wire.

FIG. 1.

3. Fill a flask completely with water, and close it with a perforated cork through which passes a narrow glass tube (open at both ends), the lower end flush with the bottom of the cork, Fig. 1. Apply a flame.

What properties change ?

4. When the water is heated nearly to boiling, remove the cork and pour a little on a lump of sugar in a saucer. Do the same with cold water. What difference is there?

5. Half fill two beakers with hot and cold water respectively. Into each pour cautiously *a few drops* of concentrated sulphuric acid. Note the difference.

6. Half fill the flask of Fig. 1 with water and heat until it boils.

What fills the upper part of the flask and tube? Is it visible? Are its properties the same as those of water? What do we see at a short distance from the end of the tube?

Hold a cold plate a little in front of the opening of the tube.

Hold the flask with a towel, invert quickly, dipping the end of the glass tube in water.

7. Put a few crystals of iodine in a test tube and warm the bottom of the tube gently over a flame.

What analogy is there between this experiment and the last?

8. Half fill a test tube with small lumps of alum; cover with water and heat to boiling. Pour off the clear liquid into another test tube, and allow to cool. Add a few drops of water, and heat again. Find the temperature at which it boils.

9. Put a pinch of powdered bluestone in the bottom of a dry test tube. Hold the test tube in a slanting position, and heat very gently over the flame of a spirit lamp. (If heated too strongly the solid turns black.)

What change do you note in the appearance of the solid? What has collected on the walls of the test tube?

When cool, add one small drop of water to the residue in the bottom of the tube. What change takes place?

10. Repeat experiment 9, using a little sawdust instead of the bluestone.

The changes which occurred in the ten experiments just performed are typical. In every case the properties changed on heating. In some cases the change was permanent, but in others the original properties were restored on cooling. The latter are called "reversible" changes. Classify the experiments from this point of view.

In works on Chemistry, the word "reaction" is often used in place of "change." The changes which take place in experiments 1 and 3 are, however, not usually called reactions, although, as the experiments show, the temperature and density of hot water are different from those of cold, as is also the amount of alum it can dissolve, the rate of its action on sugar, and indeed all the properties that can be measured quantitatively, and there is a similar difference between hot and cold platinum. This restricted use of the word "reaction" has its origin in the past history of the science of chemistry, and does not correspond to any real difference between the various "changes."

MECHANICAL MIXTURE.—EVIDENCES OF REACTION.

The experiments of the preceding section show that when two substances are brought together their properties may change, that is to say, a "new substance" may be formed. Yellow sulphur and white silver, for instance, give rise to a black substance; solid alum and liquid water form a liquid whose properties (boiling point, density, etc.) are different from those of water.

It very often happens, however, that no change takes place; each of the substances brought together retaining its properties unchanged. They are then said to form a "mechanical mixture."

Experiment:

Boil about a quarter of a litre of water in a flask or beaker and note the temperature by a thermometer.* Then add a quantity (about 50 grams) of manganese dioxide, fine sand, or chalk, and again note the temperature of the boiling liquid. Allow the substance to settle, and taste the clear liquid. Compare results with those of the experiment with alum, p. 11.

If, on the other hand, a change or "reaction" takes place when two or more substances are brought together, the product is not a "mechanical mixture" of these substances. In the following experiments there is ample evidence of change of color, temperature, or volume: in some of them the original substances may be seen to disappear, and in others, the formation of a new substance is indicated by effervescence, or explosion, or by the formation of a solid "precipitate."

After performing each of the following experiments, the pupil should give reasons for saying that the products are not mechanical mixtures of the materials brought together. They may be, of course, and very often are, mechanical mixtures of several products, or of new substances with some of the old ones left over. Write down in each case what happens and what properties change.

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* A few pieces of broken clay pipe, or unglazed earthenware, should be placed in the vessel to prevent superheating, which causes "bumping."

SUPPLIES: Solutions of copper sulphate, lead acetate, potassium iodide, and mercuric chloride; scraps of zinc, table salt, solid caustic potash, or caustic soda, sulphuric and hydrochloric acids, graduates and thermometers, mercuric oxide, and magnesium ribbon.

Experiments:

1. Into a solution of copper sulphate dip the blade of a knife, or any clean piece of iron, and allow it to remain a short time. What collects on the iron?

2. Make a solution of lead acetate in a beaker and hang in it a few strips of zinc. Set it away for a day.

Collect the solid found in the beaker and press or hammer it closely together. Cut it with a knife and examine the clean surface.

Will it make a mark on paper? What is it?

3. Measure into a beaker 60 cc. of water, and to this add 20 cc. of sulphuric acid.

When these two substances are brought together, they unite to form a homogeneous liquid whose properties (such as taste, specific gravity, boiling point, freezing point, fluidity, power to refract and reflect light, and to conduct electricity) differ from those of either of them: this alone is evidence that the product is not a "mechanical mixture."

What other evidence of a change may be noticed? Feel the beaker. Measure the volume of the liquid in the beaker.

4. To a little water in a test tube add a small piece of caustic soda, or caustic potash, and shake.

5. Put some water in a beaker, take the temperature by means of a thermometer, add a quantity of sal ammoniac (ammonium chloride), stir with the thermometer and note the temperature.

6. Place some scraps of zinc in a test tube and add some dilute hydrochloric acid.

Bring a lighted match to the mouth of the test tube.

7. Ignite a match, a scrap of paper, or a piece of magnesium ribbon.

What changes take place?

8. To a solution of potassium iodide add a solution of any lead salt, *e.g.*, lead acetate.

9. To a dilute solution of potassium iodide add slowly a dilute solution of mercuric chloride; continue until the colored substance ceases to disappear.

How do the properties of the substance formed differ from those of the substances put together?

In the preceding experiments, it is evident that there is a change of properties when the various substances are brought together, and this is sufficient evidence that they are not simply "mechanically mixed." In cases where it is difficult to decide whether a change has taken place or not, it is necessary to make a careful *quantitative* study of the properties of the supposed mixture, and of the substances from which it was formed, to see whether any change has taken place; an example will be given later in Chapter X.

Many examples of mechanical mixtures occur in nature. In a lump of granite for instance, the pieces of quartz, feldspar and mica can be seen side by side; mud is a mechanical mixture of various solids with impure water which is a solution; smoke is a mixture of solid particles and flue gas; other examples are—milk, fog, earth and many rocks.

RATE OF REACTION.

The rate or speed at which the changes took place in the experiments of the preceding section was very different in the different cases: in experiments 3, 8 and 9 the change took place very quickly, in the others it proceeded slowly. One of the circumstances on which the rate of change depends is the temperature, as is shown by the action of hot and cold water on sugar (Exp. 4, p. 10); others are illustrated by the following experiments :—

SUPPLIES: Test tubes, bluestone (crystallized copper sulphate), rock-salt, alum, marble, hydrochloric, sulphuric and oxalic acids, solutions of potassium iodide and potassium permanganate, platinum or copper wire.

Experiments :

1. Put a piece of bluestone (or rock-salt, or alum) the size of a pea, in half a test tube of water, shake, and find the number of minutes it takes to dissolve.

Repeat the experiment with the same amount of bluestone in the form of a fine powder. What circumstance has affected the rate of the reaction?

2. On a piece of marble in a test tube, pour a little very dilute hydrochloric acid; repeat the experiment, using more concentrated acid.

3. Dissolve a little oxalic acid (about 2 g.) in half a litre of water, and add about 15 or 20 cc. of concentrated* sulphuric acid. To half a test tube full of this solution, add enough potassium permanganate solution

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* By "concentrated" sulphuric acid is meant commercial sulphuric acid, to distinguish it from "dilute" sulphuric acid, the use of which is necessary in some experiments, *e. g.*, in the preparation of hydrogen.

to color the whole light red. Divide in two parts, heat one, and let the other stand. Compare the time it takes for the color to disappear in the two cases.

4. Cover a piece of the purest zinc obtainable, with dilute hydrochloric acid. Repeat the experiment, but wrap a piece of platinum or copper wire around the zinc.

Substances which affect the rate of change (or reaction), without themselves being altered, are called *catalysers*. (See pp. 27 and 28.)

5. Add hydrochloric acid to a dilute solution of potassium iodide in water, divide in two parts, using open vessels, and expose one of them to direct sunlight, while the other is kept in a dark place.

Make a list of the circumstances which have affected the rate of change in the foregoing experiments.

The above experiments illustrate some of the conditions which influence the rate at which changes take place, or the "rate of reaction," as it is usually called. In each case, the reaction took place slowly without the influence of the new condition (higher temperature, fine division of material, presence of third substance, etc.); the effect of the latter being merely to accelerate a process which would otherwise have taken place more slowly. It often happens, however, that a reaction will not occur at all, unless certain conditions prevail, although all the necessary materials are present. Examples of this will be met with in the following section, and in a review of the work the pupils should point out these conditions.

CLASSIFICATION OF REACTIONS.

In the preceding sections a large number of reactions have been met with; the present section shows how these may be divided into groups for purposes of classification.

Explosive Reactions.

Familiar examples of changes which come under this head are furnished by the behavior of gunpowder, dynamite, etc., which remain without change until the reaction is set going by a blow or by the application of heat; once begun, enough heat is liberated to keep up the necessary temperature.

Experiment:

Powder some potassium chlorate and white sugar separately and mix them on a sheet of mica. Heat over the flame of a spirit lamp, and when the action begins remove the lamp.

Such substances are called explosives. In common language the word explosion is used to denote a change which takes place suddenly with evolution of heat and sound, and with great increase in volume. From the chemical point of view, however, these explosions are very similar to the ordinary combustion of wood, coal and magnesium ribbon. Like the gunpowder, the mixture of wood and air may be kept for a long time without undergoing change; but once a spark is applied, the reaction begins and goes on until all is consumed. The rate at which the wood is burned depends on the surface exposed to the air (compare Exp. 1, p. 16). Everyone knows that shavings burn more quickly than logs, while wood dust suspended in the air may give rise to the most disastrous explosions, as those working in cork factories have discovered. Flour mills have been blown up by a similar reaction. The chemical change in all these cases is the same, and the violent mechanical effects produced when the wood is in the form of powder, are due merely to the rapidity with which large volumes of the heated gases are generated.

In the examples just considered, a small spark serves to kindle a great fire; in those which follow, a minute speck of the proper substance is all that is necessary to initiate the reaction. In both, the magnitude of the effect is out of all proportion to that of its apparent cause, and in both the explanation is the same—the reaction itself, once set going, keeps generating more and more of the substance or of the heat that was necessary for its inception.

Reactions of this type are now grouped under the general name "explosions," although they may occur without the accompaniment of light and sound, for these phenomena, as has been pointed out, depend merely on the rate at which the change takes place. From the chemical point of view the combustion of gunpowder in the fuse of a firecracker is just as much an "explosion," or "explosive reaction," as the final detonation when the spark reaches the charge, and the noiseless changes which occur in the following experiments belong to the same class.

SUPPLIES: Test tubes, beaker, spirit lamp, glass plate, sodium acetate, or sodium thiosulphate, sodium sulphate.

Experiments:

1. Put a little crystallized sodium acetate, or sodium thiosulphate in a test tube, close it with a plug of cotton

wool, and heat in a beaker of boiling water. When all is melted, remove the tube and cool under the tap, *without shaking*. Then remove the cotton, and drop in a minute speck of the acetate or thiosulphate, as the case may be.

What do you observe? Test the temperature of the tube with the hand.

To carry out this experiment successfully, it is necessary that none of the solid originally present should be left unmelted near the top of the tube. Crystallized sodium sulphate may be used in place of the acetate or thiosulphate, but in this case a white powder is formed on melting the crystals, which will not disappear on further warming.

2. The experiment may be varied, by pouring the warm solution over a clean glass plate, and, when it is cool, touching one corner with a crystal.

Many other substances behave in the same manner; even water can be cooled below its freezing point without the formation of ice, but if a crystal of the latter be added, freezing takes place at once. Compare the appearance of the plate in experiment 2 with that of a window covered with frost in the winter time.

3. About half fill a test tube with crystallized sodium sulphate, add enough water to almost cover the crystals, heat until the salt is dissolved and the liquid boils, close the tube with cotton wool as before, and allow to cool. Add a speck of the sodium sulphate crystal.

Non-Explosive Reactions.

When a spark falls on gunpowder it raises the temperature of a small portion of the powder and thus causes a reaction to take place. If the spark fell on ice it would do much the same, the reaction here consisting in the change of ice into water. There is, however, an important difference between the two cases. The change which the gunpowder undergoes (burning) is accompanied by the production of heat, which raises the temperature of adjacent particles of powder and causes the reaction to spread, while the change which the ice undergoes (melting) is accompanied by the absorption of heat, and as soon as the spark is cooled the melting ceases.

It is well known that the amount of ice melted is proportional to the amount of heat absorbed; and in other non-explosive reactions which may be brought about by the action of light, or of electricity, the amount of change has been found to be proportional to the amount of light absorbed, or to the amount of electricity passed through the solution. In fact, some of the most accurate and convenient methods of measuring heat, light and electricity are based on the measurement of chemical change.

Quantitative measurements to illustrate these relations would be out of place here, but the following experiments show the marked contrast between the reactions of this class and the explosive reactions.

SUPPLIES : Beakers, stand, spirit lamp, thermometer ; dry cells, platinum strips, or carbon pencils ; sulphuric acid, ice, solution of copper sulphate, mercuric chloride, ammonium oxalate.

Experiments :

1. Half fill a beaker with powdered ice, set it over a flame and stir with a thermometer. Remove the flame occasionally, and continue the stirring. Note that the amount of water formed depends on the amount of heat supplied, and contrast with the experiments of the previous section.

2. Connect three or four dry cells in series; to the terminal wires attach strips of platinum, or carbon pencils, and dip them into dilute sulphuric acid in a beaker.

3. Repeat the experiment with a solution of copper sulphate.

4. Dissolve a little mercuric chloride in water and add a solution of ammonium oxalate.* Divide in three portions, keep one in the dark, a second in the ordinary light of the room, and hold the third for a few minutes in direct sunlight.

Many of the reactions utilized in photography are analogous to this.

Reversible Reactions.

Another way of grouping reactions is into "reversible" and "non-reversible." When ice is heated, for instance, it turns into water; and the water may be turned back into ice by cooling it. The second reaction is the reverse of the first, and the change of ice into water is therefore said to be "reversible," or to be "a reversible reaction." The same is true of the change of water into steam, and, in general, of most cases of melting and boiling.

The changes which platinum and iodine undergo on heating (Exps. 1, p. 10, and 7, p. 11) are also obviously reversible, although, as has been explained on page 12, it is not customary to speak of the former as a "reaction" at all. When bluestone is heated (Exp. 9, p. 11) it is converted into a white powder, which remains

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* A satisfactory result may be obtained as follows: dissolve separately about one gram of each substance in about 20cc. of water, and then mix the solutions. If the mercuric chloride solution is not clear it should be filtered before mixing with the ammonium oxalate solution.

at the bottom of the test tube, and steam, which condenses on the sides. If the tube be allowed to cool, and the water be shaken down on the powder, bluestone is formed again, showing that this reaction too is reversible. So also is the reaction of alum with its solution in water. (Exp. 8, p. 11.)

Opposite changes (for example, ice into water and water into ice) are often spoken of as the same reaction taking place "in opposite directions."

An example of a reaction which takes place in one direction in the light and in the other in the dark is furnished by the following experiment.

5. Add hydrochloric acid to a very dilute solution of silver nitrate in a small stoppered bottle. Expose for a few minutes to sunlight.

What change takes place in the white solid produced?

Put away in the dark for a few days, and examine again.

The changes which occur in the leaves of plants during growth probably belong to the same class of reaction.

Non-Reversible Reactions.

When sawdust is heated (Exp. 10, p. 11), the change that takes place is said to be "non-reversible," because cooling and mixing the tar, smoke, and charcoal produced will not make sawdust again. Further examples are furnished by experiments 8, p. 8; 7, p. 15; and 4, p. 17. Means have been found to reverse many reactions which at first sight appear to be non-reversible. Examples of these will be met with later.

Chemical Equilibrium.

This term is applied to an important group of phenomena, of which the reaction between water and ice furnishes a good example.

As has already been explained, the reaction between ice and water takes place "in one direction" at high temperatures, and "in the other direction" at low temperatures. There is one intermediate temperature at which it takes place "in neither direction," that is, at which ice and water may be brought together without either melting or freezing. The ice and the water are then said to be "in equilibrium." This temperature is called the temperature of equilibrium. It changes slightly with the pressure, an increase of one atmosphere (760 mm.) lowering it less than the one-hundredth part of a degree.

The temperature of equilibrium between ice and water at atmospheric pressure has been selected as the zero of the centigrade thermometer; it is commonly called the melting point of ice.

Water and steam may likewise be in equilibrium, and the temperature of equilibrium (boiling point) at atmospheric pressure has been selected as the 100° mark on the centigrade thermometer. The boiling point of water changes with the pressure much more than the freezing point does. At two atmospheres pressure water boils at about 120° C.

Above zero (at atmospheric pressure) water cannot freeze, and below that temperature ice cannot melt; but as has been pointed out on page 20, it is quite possible to cool water below zero in the absence of ice. The water is then said to be "supercooled," and if a speck of ice be introduced an "explosive" reaction takes place. Similarly with the reaction between water and steam: above 100° C. (at atmospheric pressure) steam cannot turn into water, and below that temperature water cannot turn into steam (of atmospheric pressure). But as in the previous case, the reverse is not true. In a clean vessel water may be heated above its boiling point, or superheated, as it is called, and the "explosive" reaction that often takes place is known in the laboratory as "boiling with bumping."

Another case of equilibrium already met with is that between a solid and its saturated solution. The supersaturated solutions of Exps. 1, 2, 3, p. 20 are somewhat analogous to supercooled water.

CHAPTER II.—OXYGEN.

SUPPLIES: Test tubes, spirit lamp, perforated cork and delivery tube, mercuric oxide, potassium chlorate, manganese dioxide, ferric oxide, copper oxide, or clean sand.

Experiments:

Experiments 1 and 2 to be done by the teacher.

1. **Heating Red Oxide of Mercury.** Heat a little red oxide of mercury (about 2 grams) in a dry test tube, shaking the tube to keep the powder from sticking to the glass. Note the change of color. After heating a few minutes introduce a glowing splinter into the mouth of the tube.

What has gathered in the upper (cooler) part of the tube?

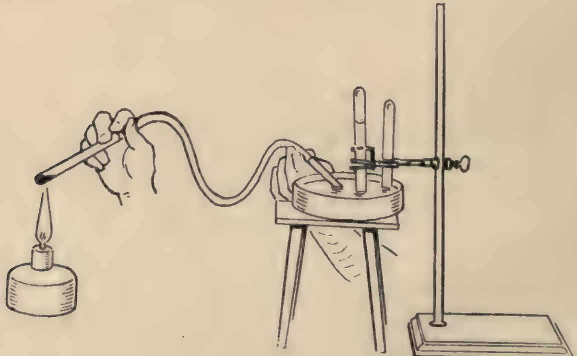

FIG. 2.

2. **Collecting the Gas.** Heat about 3 grams of the red oxide in a test tube provided with a perforated rubber cork, through which passes an inch of glass tubing con-

nected to about 18 inches of small-bore rubber tubing. Keep shaking the tube in the flame with the right hand, and when the deposit of mercury has begun to form, collect two test tubes full of the gas as shown in Fig. 2. (Do not dip the rubber tube deeper than necessary under the water, or the pressure may blow out the walls of the hot test tube. The second tube full should be collected without interrupting the heating. Why?)

Into the first tube insert a glowing splinter; close the second under water with a soft rubber stopper, and keep (with the mouth under water) for experiment 2, p. 30.

In these experiments the red powder has been converted into mercury, and a gas called oxygen. In chemical language, the red powder is said to "contain" or "be composed of" mercury and oxygen, and its name "mercuric oxide" or "red oxide of mercury" is intended to remind one of this fact.

Mercuric oxide is expensive, and the preparation of oxygen from it is inconvenient. Fortunately there are other substances which, on heating, give off that gas. One of them, which is often used for this purpose, is called "chlorate of potash." The meaning of this name will be explained later.

3. **Action of Heat on Chlorate of Potash.** Heat a little of this substance (about 2 grams) in a test tube. When the melted substance begins to bubble, hold a glowing splinter in the tube. Heat until no more oxygen is given off. Is anything left in the tube?

4. **Catalytic action of Manganese Dioxide, etc.** Heat some of the chlorate as before. When it begins to give off gas, drop in a little finely powdered manganese dioxide from the point of a knife. Note the change in the rate at which the gas is given off.

Try the same experiment in another tube, using ferric oxide, copper oxide, or clean sand, instead of manganese dioxide.

No satisfactory explanation of the action of the manganese dioxide (or the ferric oxide, etc.) in this experiment has been found. As may be shown by simple comparative experiments, the amount of oxygen that can be obtained from a given weight of chlorate is the same whether these other substances be added or not. They themselves are not used up, and do not seem to suffer any change by being heated with the chlorate. If the contents of the tube (Exp. 4), after cooling, be ground up with water, the dioxide remains undissolved, and may be collected on a filter and used again and again.*

Substances like these, which affect the rate of a chemical change, without apparently undergoing any change themselves, are called "catalytic agents" or "catalysers." We have already had examples of this in experiment 4, p. 17. Others will be met with later.

The preceding experiments lead to a good method for the preparation of oxygen.

PREPARATION AND PROPERTIES OF OXYGEN.

SUPPLIES: Large test tube, cork and delivery tube, bottles for collecting gas, splinters, combustion cups, or deflagrating spoons, potassium chlorate, manganese dioxide, pyrogallic acid and caustic potash, sulphur, phosphorus, potassium or sodium, magnesium ribbon, picture wire.

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* If this be attempted, care should be taken not to introduce any shreds of filter paper into the mixture with the chlorate; see note, page 29.

1. **Preparation of Oxygen.** Mix potassium chlorate and manganese dioxide[*] in about equal proportions, and half fill a large test tube with the mixture. Fit the test tube with a cork and delivery tube, which leads to the mouth of a bottle[†] filled with water, and inverted in a trough also filled with water. Arrange the apparatus as shown in Fig. 3, and heat the test tube with a low flame.

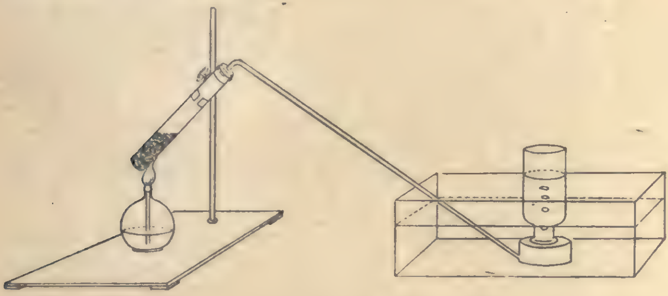

FIG. 3.

Do not collect the first gas that comes off. Why?

If the gas is evolved too rapidly, heat less strongly. When the gas ceases to come off, disconnect the delivery tube. Why?

Collect as many jars of gas as possible, and allow some to stand in the water till next day.

Make a list of as many properties of oxygen as may be noted, and compare them with those of air.

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[*] Manganese dioxide is sometimes mixed with powdered charcoal, and is then very dangerous to use; a small sample of the mixture of chlorate and dioxide should be heated in a test tube before it is given to the class. If it explodes, the dioxide is impure and should not be used. Even when pure, the chlorate and the dioxide should be powdered separately and mixed carefully on a piece of paper. They should on no account be ground together in a mortar.

[†] An ordinary pickle bottle will serve very well; fill it with water, cover with a glass plate, or a piece of paper and invert it in the trough.

Tests for Oxygen.

The gas thus prepared from potassium chlorate may now be compared with that from mercuric oxide in appearance and in its action on a glowing splinter. The test tube full of gas kept from experiment 2, p. 26, and another filled in experiment 1 above, may be used in the following experiment :—

2. **Action of Pyrogallic Acid on Oxygen.** Fill a test tube with oxygen, and, holding the finger tightly over the open end, place the mouth of the tube below the surface of a solution of caustic potash (potassium hydroxide) and pyrogallic acid in a beaker. Remove the finger and shake the tube for some *minutes*, being careful to keep the mouth of the tube below the surface of the solution.

By its appearance alone, oxygen could not be distinguished from air or coal gas, but its behavior with a glowing splinter and with pyrogallic acid is different from that of any other gas, and may be used to detect or identify it. One may "test" for oxygen with these two "reagents for oxygen." There are, of course, many others. One test is seldom sufficient to distinguish a substance from *all* others, and to be quite sure of the identity of a material it is generally necessary to apply several tests.

Combustions in Oxygen.

Experiments :

3. **Oxygen and Wood.** Slip a glass plate, or a piece of paper, under the mouth of one of the bottles of oxygen ; remove the bottle from the water, and insert a glowing splinter of wood. Remove the splinter if there is any change, and repeat as many times as possible.

Describe the result. Does the oxygen burn?

4. **Sulphur and Oxygen.** Put some sulphur into a deflagrating spoon, or into a combustion cup,* Figs. 4 and 5. Heat in the flame of the spirit lamp until the sulphur catches fire, and then lower it into a bottle of oxygen. Cover the top of the bottle with a piece of glass or paper, and note carefully what occurs.

Describe any change in the flame. Smell the vapor cautiously by wafting a little toward you. What does it smell like?

Pour some water into the bottle containing the products of combustion of sulphur in oxygen, then place the hand tightly over the mouth and shake vigorously for some time.

Describe what occurs. Why is the hand sucked into the mouth of the bottle?

Drop into the bottle small pieces of red and blue litmus paper.

What does the result indicate?

To answer this question, partly fill two beakers with water; to one of them add a few drops of sulphuric, nitric, hydrochloric, or acetic acid, and to the other a little caustic potash, caustic soda, or ammonia solution. Into each put small pieces of red and blue litmus paper, and change them about from one beaker to the other. In the beaker containing the acid, the blue litmus changes to red, and in the other, which contains a "base," the reverse change

FIGS. 4 AND 5.

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* A combustion cup may be made from a piece of crayon. Make a small cavity in one end and wrap a piece of copper wire around the crayon as in Fig. 5.

takes place. This is a method often used to distinguish between acids and bases.

What change took place in the litmus of experiment 4? Was the liquid acid or basic?

5. **Phosphorus and Oxygen.*** Clean the deflagrating spoon, or make another combustion cup, and in it place a piece of phosphorus. Ignite it, and quickly lower it into a bottle of oxygen.

Describe what takes place.

6. **Potassium, or Sodium, and Oxygen.†** Place a piece of potassium about the size of a pea in a combustion cup, ignite it, and lower it into another bottle of the gas. Cover as before. Note carefully what takes place.

Add a little water, shake well as before, and then test with litmus paper.

Is the liquid acid or basic?

Evaporate some of the liquid in a dish.

What is the appearance of the residue? Is it potassium?

7. **Magnesium and Oxygen.** Take one end of a piece of magnesium ribbon in a pair of forceps, light the other end and hold it in a jar of oxygen.

In what ways does the result resemble those of the other experiments?

Add water to the powder in the bottle, shake well, and test with litmus paper.

Is the liquid an acid or a base?

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* This experiment should not be performed by the class, as phosphorus is much too dangerous a substance to be handled by the pupils. It ignites very easily, burns violently, and spurts in every direction, inflicting burns which are very painful and take a long time to heal. *Phosphorus must be kept and cut under water*, and must never be taken in the fingers even when wet.

† Potassium and sodium are other substances that should be handled with great care. A drop of water coming in contact with either may cause a dangerous *explosion*, and if touched by wet fingers painful wounds are inflicted.

8. Iron and Oxygen. Spread out the end of a piece of picture wire, heat it and dip it in sulphur. Ignite the sulphur that adheres, and hold the end of the wire in a jar of oxygen.

If a suitable piece of wire be found, it will burn, throwing off bright sparks, while the product of combustion falls to the bottom in the form of red hot iron oxide. To prevent the glass being cracked, it is well to pour some water into the bottle before performing the experiment.

The oxide of iron is insoluble in water, but the fumes from the burning sulphur would cause the liquid to give an acid reaction, if tested with litmus.

Experiments 3 to 8 should be repeated with air instead of oxygen. Mention all points of similarity and difference in the two sets of experiments. Point out any similarities in the products formed from the oxygen and the air.

Oxygen in Air.

9. **Pyrogallic Acid and Air.** Into a graduated tube, such as a eudiometer, the total volume of which is known, put some pyrogallic acid (about as much as can be held on a cent), and add about 20 cc. of strong solution of caustic potash; cork tightly and read the volume. Shake well for some time, place the mouth of the tube under water and remove the cork. Depress the tube until the water stands at the same height inside as out, and take the reading again.

What gas is dissolved by the pyrogallic acid solution? Test the residue with a splinter. Does it support combustion?

This experiment shows that air contains a substance with properties different from those of oxygen. Roughly, what proportion of the total volume does this substance fill?

Note. Air contains oxygen and another gas called nitrogen, which will be described later. It is owing to the dilution of the oxygen by this gas that the rate of the combustions in air was less than in pure oxygen.

REVIEW OF THE EXPERIMENTS WITH OXYGEN.

The experiments in the preceding section furnish a number of further examples of "reactions." (See pp. 10 to 15.)

Some of these reactions occur on simply bringing two substances into contact, pyrogallate of potash and oxygen for instance; in other cases heat is necessary, either to start the reaction (sulphur and oxygen), or throughout its whole course, as in the preparation of oxygen. The experiments on the preparation of oxygen also illustrate the effect of temperature and catalytic agents on the rate of a reaction.

Chemical Terms.

With the exception of the names of the chemicals, the only strange word made use of is "catalytic," explained on page 28. A number of common English words, however, have been used in a special sense, and thus form a body of technical terms whose meaning should be carefully noted. Thus "reaction" in chemistry has much the same meaning as "action" in common language; "preparation" means manufacture, and "combustion" burning. The meaning of "slow combustion" will be explained in the next section.

We have had examples of materials which, without the permanent addition of any further matter, have been changed into two or more entirely distinct substances, *e.g.*, potassium chlorate gave oxygen and potassium chloride ; mercuric oxide gave mercury and oxygen ; a solution was converted into a solid and water, etc. There are, however, a comparatively small number of substances which all attempts have failed to decompose in this way: These are called "elements." Oxygen belongs to this class, and there are about 80 others known. A complete table of the elements will be found on page 215.

It should be pointed out that if a substance can be formed by the combination of two or more other substances quantitatively, this fact alone is sufficient to show that the substance in question is not an element.

The red powder which, on heating, turns into mercury and oxygen is said to "contain" mercury and oxygen, although it has none of the properties of these elements, and similarly, the gas formed by burning sulphur in oxygen is said to "contain" sulphur and oxygen. Potassium chlorate, which on heating turns into oxygen and a white powder (potassium chloride), "contains" oxygen, but it is not said to contain potassium chloride, because it is usual to speak only of elements as being contained in such substances. There are, however, exceptions to this custom, which will be met with later.

Chemical Classification and Nomenclature.

Chemicals are classified by their composition (that is, according to what they "contain"), and by their reactions.

Potassium chlorate and mercuric oxide, for instance, are "compounds of oxygen," and those compounds of oxygen which contain only oxygen and *one* other "element" are termed "oxides." When there is more than one oxide of an element, they are distinguished by an adjective—"black oxide" and "red oxide" of mercury, for instance, or by modifying the form of the words used to tell what they contain, as "mercurous oxide" and "mercuric oxide." The systematic use of these endings will be explained later.

An example of classification by reactions, is furnished by the groups "acids" and "bases"; any substance which turns blue litmus red being classed as an acid, while those which turn red litmus blue are called bases.

In the experiments described on pp. 31–32, sulphur and phosphorus, on burning in oxygen, gave oxides which with water produced solutions with acid properties, while the oxides of potassium and magnesium gave basic solutions. This behavior is the basis of a rough classification of the elements into acid-forming and base-forming elements. It corresponds very closely with a classification based on the general appearance of the elements, into non-metals and metals.

To which class does sodium belong, judged by its appearance? How would you settle whether it was an acid-forming or a base-forming element?

NOTES ON OXYGEN.

Oxygen was first discovered by Priestly in the year 1774, by heating mercuric oxide. It is a colorless, odorless, tasteless gas, only slightly soluble in water, 100 cc. of water at 0°C. dissolving about 4 cc. of oxygen. Its density, as compared with air, is 1.1056 and 1 litre at

0° C. and 760 mm. pressure weighs 1.429 grams. At ordinary temperatures, oxygen does not act readily upon most other elements, but if heated it unites rapidly with the majority of them, forming oxides.

For a long time oxygen was known only in a gaseous state, but in 1877 it was converted into a liquid. This has a bluish color and boils under atmospheric pressure at −180° C.

The preparation of oxygen from the air has, since 1896, become of commercial importance. The method employed is explained in the chapter on air (see p. 161).

COMBUSTIBLES, and SUPPORTERS OF COMBUSTION.

By a "combustible" substance is ordinarily meant one that will burn in oxygen (or in air), while the oxygen itself is termed a "supporter of combustion." That this is a distinction without a difference may be shown by the following experiments.

Experiments :

1. **Combustion of Oxygen in Alcohol Vapor.** Spread out the top of the wick of a spirit lamp, so as to produce a large hollow flame when lighted. Into the centre of this flame, which consists of alcohol vapor, introduce a stream of oxygen through a glass tube drawn to a fine point.

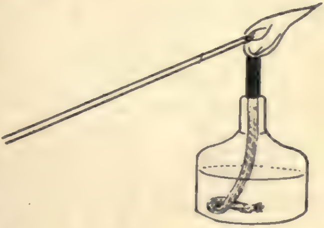

FIG. 6.

When we consider the flame of the lamp under ordinary conditions, what is regarded as the combustible substance, and what as the supporter of combustion?

When the current of oxygen is introduced into the centre of the flame, what is seen to be burning? What is now the combustible substance, and what the supporter of combustion?

2. **Combustion of Oxygen in Illuminating Gas.** Cover the upper end of an ordinary lamp glass with a cap of fine wire gauze, and close the bottom with a cork, through which pass two tubes, as in Fig. 7. One of these tubes is about an inch long, straight, and of larger diameter than the other. Attach the bent tube to a supply of illuminating gas; let it flow for a time, until the chimney is filled. Light the gas at the top of the chimney, and turn it down until the flame is about five inches high. Then thrust a lighted match into the lower tube.

FIG. 7.

We have here the phenomena of coal gas burning in air at the top of the chimney, and air burning in an atmosphere of coal gas at the top of the short tube. As in the previous experiment, what in one case is regarded as the supporter of combustion must be spoken of in the other as the combustible substance, and *vice versa*.

Since it is thus possible, by varying the conditions, to render these terms interchangeable, it might be better to express the relation between such substances by saying that they are "mutually combustible."

TEMPERATURE OF IGNITION.

A substance when heated will "ignite," or catch fire, only if the heat applied raises the adjacent parts of the substance to a temperature at which the chemical action takes place quickly enough to produce sufficient heat to keep the temperature from falling. The "temperature of ignition," therefore, depends not only on the chemical nature of the substance, but also on its power to conduct or radiate heat.

Experiments :

1. **Ignition of Carbon Disulphide.** Put a few drops of carbon disulphide in a test tube and shake well. Smell cautiously. Heat the end of an iron wire and bring it to the mouth of the tube.

What does the experiment show? Why is it more difficult to light a coal fire than a wood fire? Which has the lower temperature of ignition, sulphur or phosphorus?

2. **Ignition of Iron.** Sprinkle some fine iron filings in the flame of a spirit lamp.

Hold in the flame a piece of iron wire gauze.

Set fire to a small heap of "ferrum redactum" with a match.

The small iron specks get hot quickly, and lose none of their heat by conduction, while the gauze conducts the heat away, and gives it off to the air.

When combustion takes place in the air, not only must the substances themselves and the products of combustion be raised to a high temperature by the heat produced, but the nitrogen, which makes up about four-fifths of the volume of the air, must likewise be heated; this keeps down the temperature of the flame. We have observed that when substances burn in oxygen, the light emitted is much more intense than when the same substances burn in air; this is due to the higher temperature of the flame in oxygen: indeed, one of the methods commonly employed for determining the temperature of incandescent bodies is based on measurements of the light given off.

Picture wire burns in oxygen; why not in air?

SLOW OXIDATION.

Many substances combine slowly with oxygen without, apparently, any heat being evolved. If iron be allowed to lie in moist air it becomes covered with rust, and if potassium be cut, the surface which at first is bright, soon changes to a dull grey, owing to the action of the oxygen of the air. As a matter of fact, heat is produced in these reactions, and the total quantity of heat produced, when a given weight of substance is changed, is the same whether the change be fast or slow. But when the action is rapid, the temperature rises, perhaps even to the ignition point; while if it is slow, the heat may be conducted away as fast as it is generated, and the rise in temperature may be so small as to be inappreciable.

A very important instance of slow oxidation is that which takes place within our own bodies. The food which we eat undergoes many changes, in which oxidation plays an important part. We take large quantities of oxygen into our lungs when breathing, and some of this is absorbed by the blood and carried by it to all parts of the body, where it reacts with many of the substances presented to it, oxidizing them to form products which are afterwards eliminated from the system. This process of oxidation is accompanied by the production of the heat that keeps our bodies warm.

If the heat generated by slow oxidation cannot escape, the temperature rises, and if the ignition point be reached, we have "spontaneous combustion." Piles of soft coal, especially kinds that contain much sulphur, furnish an example; and oily rags or waste lying in a heap, frequently take fire from the same cause. Instances are not unknown where barns have been burned because of being filled with damp grain, or because of the oxidation of piles of manure or damp straw lying against them.

CHAPTER III.—CONSERVATION OF MASS

CHEMISTRY is a *quantitative* science: its systems of classification, its formulae, and nearly all its important laws are based on relations between the *weights* of substances which enter into reaction. The most general of all these laws, and one on which all modern chemistry is based, is the subject of the present chapter.

SUPPLIES: Test tubes, beaker, balance; solutions of ferric chloride, potassium or ammonium sulphocyanate, potassium iodide, lead acetate, mercuric chloride; common salt, baking soda, hydrochloric acid, ferrum redactum, phosphorus, porcelain dish, lime water, a candle.

Experiments:

1. **Change of Color.** Pour a little dilute ferric chloride solution into a beaker, and partly fill a test tube with a very dilute solution of ammonium (or potassium) sulphocyanate; stand the test tube in the beaker (Fig. 8), and poise carefully on a balance. Then empty the contents of the tube into the beaker, replace the tube, and balance again.

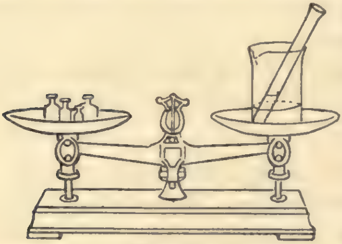

FIG. 8.

What change occurred? Was there a change of weight?

2. **Precipitation.** Repeat the above experiment, using solutions of potassium iodide and lead acetate, or solutions of potassium iodide and mercuric chloride.

3. **Solution.** Put water in the beaker, and table salt in the tube. After balancing, pour the salt into the water, and stir with the tube.

Loss of Weight during a Reaction.

4. **Baking Soda and Acid.** Using the apparatus of Exp. 1, put a little baking soda (sodium bicarbonate) in the beaker, and hydrochloric acid in the tube.

Is the weight after mixing the same as before?

When the acid was poured on the soda, there was a violent effervescence, caused by the formation of a gas, which might be collected by using the apparatus described on page 29. If this gas, or the air it displaces from the beaker, has any appreciable weight, part at least of the loss of weight would be accounted for.

5. **Has Air Weight?** Fit a round quart bottle with a glass tap (or brass pet-cock), by means of a perforated rubber stopper well tied in. Counterpoise it carefully on a balance capable of showing a good deflection with one gram additional weight. Remove from the balance, and pump it full of air by means of a bicycle pump or rubber syringe; close the tap, and replace on the balance. Note the deflection; then open the tap without removing the bottle from the pan.

6. **Baking Soda and Acid in a Closed Vessel.** In view of the result of Exp. 5, it is worth while to repeat Exp. 4 under such conditions that the gas formed cannot escape from the vessel in which the products of the reaction are weighed. This may be accomplished as follows:—

Put four grams* of baking soda in the quart bottle used in the last experiment; partly fill a test tube with

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* The gas generated by this weight of soda will not create a dangerous pressure.

hydrochloric acid and stand it in the bottle without spilling any of the acid on the soda. Then insert the stopper, close the tap, and tie both in. After counterpoising on the balance, tilt the bottle so that the acid flows out of the tube onto the soda, and balance again.

Is there any change of weight?

Open the tap while the bottle remains on the balance. Has the weight changed? How may the loss of weight in Exp. 4 be accounted for?

Gain of Weight during a Reaction.

It hardly needs a laboratory experiment to show that as a candle burns down its weight decreases, or that the ashes left from a ton of coal weigh less than the coal itself, but cases in which combustion is accompanied by an *increase* in weight are not so familiar.

7. Gain of Weight on Combustion of Iron. Counterbalance a porcelain dish containing a little heap of ferrum redactum, and then set fire to it with a match. Does the weight alter?

If the iron powder be attached to a horseshoe magnet suspended from the beam of the balance, the combustion takes place more rapidly, on account of the greater surface exposed to the air. A dish must be placed on the pan to catch any particles that may fall off.

Chemical examination of the product of the combustion shows that it is identical with that formed when iron is burned in oxygen (Exp. 8, p. 33); and the experiments with soda and acid suggest the advisability of repeating the combustion in a vessel shut off from the atmosphere.

8. Combustion of Iron in a Closed Vessel. It can be shown, although the experiment itself is not easy to carry out without the aid of an electric current, that when iron (ferrum redactum) is allowed to burn in a closed vessel containing oxygen or air, no change in weight occurs. Full directions for this experiment are given in Miller's "New Requirements in Chemistry," page 3.

Other Combustions.

9. **Combustion of Phosphorus.** In a wide mouthed flask of about 1 litre contents, place a porcelain dish containing a small piece of phosphorus (Fig. 9). Cork tightly and weigh. Then heat the bottom of the flask until the phosphorus ignites, and, after all action has ceased, weigh again.*


FIG. 9.

10. **Combustion of a Candle.** Light a candle, and bring mouth downwards over the flame a cool, dry beaker.

What collects on the inside of the beaker?

Does the same thing happen when the beaker is brought mouth downwards over a candle that is not burning?

Wet the inside of a beaker with lime water, and hold it mouth downwards over the flame of a candle. What happens?

Do these experiments suggest any explanation of the loss of weight of the burning candle?

"Law" of the Conservation of Mass.

The above experiments show that in each of the reactions studied, the weight of the substances formed (products of the reaction) is just the same as that of the

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* The flask must be cool when weighed the second time, or the heated air around it, in rising, will decrease the apparent weight of the flask.

substances used up (substances entering into the reaction). The two apparent exceptions proved on examination to be due to the evolution or absorption of gases, which although invisible, yet have weight; and when these two reactions were carried out in closed vessels, the weight of the whole was found to be the same after the reaction as before it.

This makes it seem probable that the loss of weight on burning a candle may be due to the same cause, the more so as gases evidently take part in the reaction, for without air the candle will not burn, and the presence of a gaseous product of combustion was shown by Exp. 10. This probability can be changed into certainty by burning the candle in a large closed flask, when it will be found that the weight of the flask and its contents remains unchanged.

A large number of other reactions, of the most varied types—involving change of color, precipitation, solution, combustion, etc., have been studied in the same way; and in every case, without exception, it has been found that the substances formed by the reaction weigh as much as those used up. In dealing with any particular reaction in which we happen to be interested, therefore, it is safe to assume that if it were carried out in a closed vessel on the pan of a balance, no change of weight would be observed. The probabilities in favor of the truth of this assumption are so great that, in most cases, it would appear to be a sheer waste of time to undertake special experiments to make sure.

The general statement, or "generalization," that the substances formed in a reaction weigh as much as those used up, is commonly called a "law," though it might much better be called a "probability," and the process of discovering such probabilities is termed "induction."

For very good reasons, which do not concern us, the "law" just enunciated has received the name, "Law of the Conservation of Mass"; it is the most important and the most general of the quantitative laws of chemistry.

Applications of the Law of the Conservation of Mass.

It is a very difficult matter to weigh gases accurately, but by means of the law of the conservation of mass, it is possible to find quite easily the weight of oxygen given off by a gram of potassium chlorate.

11. **Weight of Oxygen given off by one gram of Potassium Chlorate.*** Weigh out accurately in a dry porcelain crucible, about a gram of finely powdered potassium chlorate. This may be done as follows: First weigh the crucible, then add one gram to the weights, and pour the chlorate into the crucible until a little overbalanced; lastly, add weights until perfectly balanced.

Suppose the weight of the empty crucible to be	10.56 g.
And that of the crucible and chlorate	11.72 g.
Then the chlorate must weigh ..	1.16 g.

Now support the crucible on a pipeclay triangle, and heat (gently at first) until no more gas is given off; let cool, and weigh again.

Suppose that the crucible and residue weigh	11.27 g.
But the crucible and chlorate weighed ..	11.72 g.
So that the loss on heating is ..	0.45 g.

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* If some of the pupils add manganese dioxide, they can verify the statement made on p. 28, that the amount of oxygen given off per gram of chlorate is unaffected by the addition of the dioxide.

Since the chlorate on heating turns into oxygen, and the non-volatile potassium chloride, *this loss must be the weight of the oxygen given off during the heating*.

Thus, 1.16 g. chlorate gave off (or "contained") 0.45 g. oxygen, that is 1 g. chlorate contains 0.39 g. oxygen.

The application of the law of the conservation of mass is contained in the sentence printed in italics. Its truth might be tested (though not in a school laboratory) by heating the chlorate in a suitable vessel, collecting the oxygen, and weighing it.

By a slight modification of the above experiment, the *density*, or the weight of a unit volume of oxygen, may be determined with a fair degree of accuracy. If the gas be collected, as was done in experiment 1, p. 29, the volume of gas produced from a known weight of chlorate can be determined in the following way: After the evolution of gas has ceased and the apparatus has been allowed to cool, cover the mouth of the bottle with a glass plate, invert carefully, without losing any of the water, and determine the volume of the additional water necessary to completely fill the bottle. In order to prevent the water being sucked back into the test tube as the apparatus cools, the end of the delivery tube should reach well above the surface of the liquid in the bottle. The weight of the oxygen evolved from the potassium chlorate being known from the previous experiment, it is a simple matter to calculate the weight of one litre of the gas, at the temperature and pressure at which it was collected. The corresponding value for 0° C. and 760 mm. may be calculated from the gas laws. (See p. 52.)

A suitable quantity of potassium chlorate to take for this experiment is about 0.6 gram. This will produce about 200 cc. of oxygen. A little manganese dioxide should be mixed with the chlorate after weighing. The result that has been determined by more accurate methods is given on page 36.

PROBLEMS.

5.55 grams of mercury are formed by heating 6 grams of mercuric oxide. How much oxygen does one gram of the oxide contain?

How much oxygen is combined with one gram of mercury in the red oxide?

What per cent.* of oxygen is contained in the red oxide?

For the purpose of solving problems such as these, it is convenient to express the law of conservation of mass in the form of an equation, thus:—

Weight of oxide used up = weight of mercury formed + weight of oxygen formed.

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* The weight of oxygen contained in 100 grams of the oxide is called the per cent. of oxygen in the oxide.

CHAPTER IV.—OZONE

SUPPLIES : Stick of phosphorus, filter paper, solution of potassium iodide, red and blue litmus paper, electric cells, induction coil.

Experiments :

1. Place a clean stick of phosphorus in a bottle and about half cover it with water. Cork tightly and leave until next day.

Smell the gas contained in the bottle.

Hang in the bottle (*a*) a piece of filter paper moistened with a solution of potassium iodide, and (*b*) some moistened red and blue litmus paper; cork the bottle again and leave for a time.

Describe what has taken place.

2. Connect an electric cell, or battery, to an induction coil and bring the terminals of the secondary close enough to allow the sparks to pass rapidly.

Note the odor in the air near the sparks.

Place close to the spark, a paper moistened with a solution of potassium iodide. What is the result?

That something different from ordinary air is formed in the above experiments is shown by the peculiar odor and the brown coloration of the paper moistened with the solution of potassium iodide, owing to the liberation of iodine.

The same results might have been obtained with pure oxygen instead of air.

The smell is due to a new substance called ozone. This material may be prepared in a pure state and condensed to a blue liquid, which upon heating changes, weight for weight, into pure oxygen. This fact shows that the relationship between oxygen and ozone is similar to that between water and ice. In each case there are two entirely different substances having the same composition. Many other examples of this will be met with later.

Properties of Ozone.

Pure ozone is very difficult to prepare, but a small proportion of the oxygen in the air may be easily converted into ozone by such substances as phosphorus, or by electric sparks.

Ozone is found in the air of the country, especially in sea air, and in the neighborhood of pine forests. It is said to be more plentiful after electric storms. Its presence in these particular places in larger quantities may be accounted for by the facts that ozone is always formed when water evaporates, and that turpentine and certain essential oils, when acted on by air or oxygen, transform a portion of the oxygen into ozone.

Accurate experiments have established the following relations between the oxygen that disappears and the ozone formed, namely, three volumes of oxygen become two volumes of ozone, and conversely, two volumes of ozone decompose on heating, forming three volumes of oxygen.

In general ozone reacts more readily with other substances than oxygen does.

CHAPTER V.—GAS LAWS

In order to solve many chemical problems, a knowledge of the relation between the volume, temperature and pressure of a gas is necessary. Fortunately this relation is simple and is the same for all gases, so that calculations involving changes of volume, temperature, and pressure of gases are not difficult.

The volume of a fixed weight of gas is *proportional to the absolute temperature* (*i.e.*, the centigrade temperature + 273) and is *inversely proportional to the pressure.*

Consequently, (1) if the volume of a certain quantity of gas is known at any given pressure and temperature, the volume at any other pressure and temperature can be calculated.

(2) If the pressure is known for a given volume and temperature, it can be calculated for any other volume and temperature.

(3) If the temperature at which a certain weight of gas occupies a given volume, at a given pressure, is known, we can calculate the temperature corresponding to any other volume and pressure.

It need hardly be stated that two litres of a gas weigh twice as much as one litre, when measured at the same pressure and temperature, *i.e.*, the weight is proportional to the volume. It is perhaps not so obvious that the weight of one litre of a gas at 800 mm. is just twice that of one litre of the same gas at the same temperature, at a pressure of 400 mm.; in other words, volume and temperature being the same, the weight of gas is proportional to the pressure.

This will be illustrated by the following examples,—in solving which the pupil will find it helpful to keep in mind the special case of a rubber ball full of air. If the ball is *warmed*, it swells and gets harder (*increase* in *volume* and *pressure*). If the ball is squeezed (*increase* in *pressure*) its size is decreased (*decrease* in *volume*).

1. (*a*) A certain weight of oxygen fills 50 litres at 0° C. and 760 mm. pressure. Find what volume it will fill at 18° C. and 720 mm. pressure.

Solution :

0° C. corresponds to 273 + 0 = 273° Absolute,
18° C. " " 273 + 18 = 291° "
At 273° Abs. and 760 mm. pressure, the volume is 50 litres;
" 291° " " 760 " " " " " $\frac{50 \times 291}{273}$ "
" 291° Abs. and 720 mm. pressure, the volume is $\frac{50 \times 291 \times 760}{273 \times 720}$ "
= 56.258 litres.

(*b*) A certain weight of oxygen fills 50 litres at −27° C. and 720 mm. pressure. Find what volume it will fill at 45° C. and 780 mm. pressure.

Solution :

−27° C. corresponds to 273 − 27 = 246° Abs.
45° C. " " 273 + 45 = 318° "
At 246° Abs. and 720 mm. pressure, the volume is 50 litres;
" 318° " " 720 " " " " " $\frac{50 \times 318}{246}$ "
" 318° Abs. and 780 mm. pressure, the volume is $\frac{50 \times 318 \times 720}{246 \times 780}$ "
= 59.662 litres.

2. The pressure of a certain weight of oxygen occupying 25 litres, at a temperature of 15° C., is 750 mm.

Find what the pressure would be if the volume were changed to 20 litres and the temperature to 27° C.

Solution :

When the gas fills 25 litres at 288° Abs. the pressure is 750 mm.
" " " " 20 " " 288° " " " " $\frac{750 \times 25}{20}$ mm.
When the gas fills 20 litres at 300° Abs. the pressure is $\frac{750 \times 25 \times 300}{20 \times 288}$ mm.
$= 976.56$ mm.

3. A certain weight of air filling 100 cc. at 15° C. is under a pressure of 740 mm. What temperature will be necessary in order that the pressure shall be 1280 mm. when the volume is 75 cc.?

Solution :

At 740 mm. pressure the air fills 100 cc. at 288 ° Abs.
" 1280 " " " " " 100 cc. " $\frac{288 \times 1280}{740}$
" 1280 " " " " " 75 cc. " $\frac{288 \times 1280 \times 75}{740 \times 100}$
$= 373.62°$ Abs.
$= 100.62°$ C.

PROBLEMS.

1. A certain weight of oxygen fills 500 cc. at 40° C. and 730 mm. pressure. What will be its volume at 100° C. and 770 mm. pressure?

2. The pressure of a given amount of air, occupying 1 litre at a temperature of $-10°$ C., is 750 mm. What will be the pressure if the volume be changed to 1.2 litres and the temperature to 15° C.?

3. A given quantity of oxygen filling 22.4 litres at 0° C. is under a pressure of 760 mm. What temperature

will be necessary, in order that the pressure shall be 730 mm. when the volume is 25 litres?

4. 1 litre of oxygen at 760 mm. pressure and 0° C. weighs 1.429 g. What will be the weight of 25 litres at 750 mm. and 10° C.?

5. How many grams of oxygen may be obtained from 100 grams of potassium chlorate? (Make use of the data given in the experiment 10, p. 47, to solve this problem.) How many litres would this weight of oxygen fill, at 760 mm. and 0° C.? (Use data given in problem 4 for this.) How many litres would it fill at 17° C. and 760 mm.?

6. What is the weight of (*a*) 20 l. of oxygen, (*b*) 5 l., (*c*) 500 l., at 0° C. and 760 mm.? (Use data of problem 4.)

A cylindrical vessel 14 dcm. in diameter and 4 m. high is filled with oxygen at 0° C. and 760 mm. What is the weight of the gas? (Use data of problem 4.)

7. What weight and what volume of oxygen may be obtained from 500 grams of a sample of potassium chlorate which contains 15 per cent. of impurities? (Use data of Exp. 10, p. 47, or result of problem 5.)

8. What volume of oxygen at 27° C. and 800 mm. may be obtained from 612.5 grams of potassium chlorate? (See last problem.)

9. What would be the volume in the last question if it were measured at -27° C. and 670 mm.?

10. How much potassium chlorate must be heated to give 23 l. of oxygen at 14° C. and 836 mm.? (Use data of Exp. 10, p. 47, or result of problem 5.)

11. What volume of liquid oxygen, sp. gr. 0.89, would be obtained from 100 l. of the gas measured at 20° C. and 760 mm.? (Use data of problem 4.)

CHAPTER VI.—HYDROGEN

Preparation of Hydrogen. Put a quantity of zinc clippings, or granulated zinc,* into a bottle, fit it tightly with a perforated rubber stopper, through which is passed a thistle tube and a delivery tube.

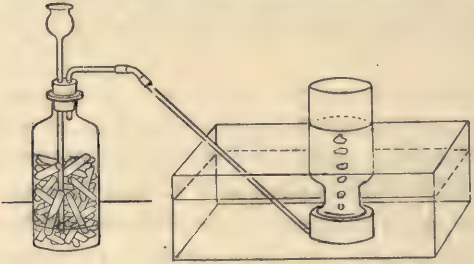

FIG. 10.

Pour down the thistle tube a small quantity of dilute hydrochloric acid or dilute sulphuric acid.† Allow the gas to escape for a short time, and then hold over the mouth of the delivery tube a test tube filled with water; when filled with gas, hold it mouth downwards over the flame of a spirit lamp. Repeat this until the gas collected burns quietly.‡ When it burns without explosion collect the gas in bottles filled with water and inverted, as in Fig. 10. (If the zinc be pure, the evolution of gas will be slow. To remedy this, add

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* Granulated zinc is made by melting zinc and pouring it slowly into cold water, from a height of four or five feet.

† Ordinary commercial hydrochloric acid may be used, but the sulphuric acid must be diluted with water in the proportion of about 1:6.

‡ On account of the danger of an explosion, this test must never be omitted when taking hydrogen from a generator, or gas-holder.

a little copper sulphate solution. See Exp. 4, p. 17, and Exp. 4, p. 27, for the reason.)

Collect several bottles of the gas and allow them to stand in water until required, being careful to allow no air to enter. When more hydrogen is required add more acid through the thistle tube.

What change is noticed in the zinc?

What properties of the gas may be noted?

Compare these with the properties of air and oxygen.

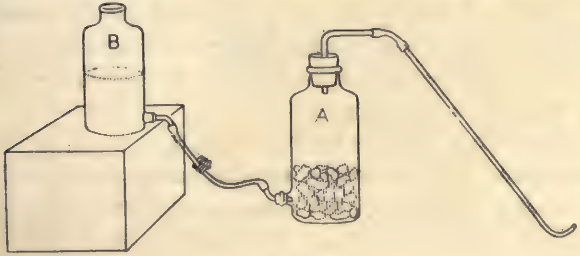

FIG. 11.

Figs. 11 and 12 illustrate other forms of apparatus that may be used for giving a supply of hydrogen as required. In Fig. 11, B is a bottle provided with a tubulure at the bottom, and containing dilute hydrochloric, or dilute sulphuric acid; A is a similar bottle containing pieces of zinc. They are connected by a rubber tube on which is a clamp for controlling the flow of acid into A. Fig. 12 is the apparatus ordinarily used for generating hydrogen sulphide. It is known as the Kipp apparatus.

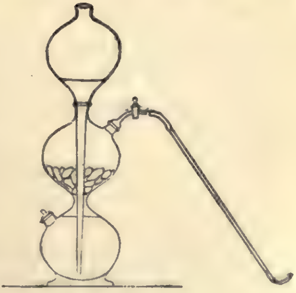

FIG. 12.

Properties of Hydrogen.

SUPPLIES: Bottles of gas freshly prepared, or carefully kept over water, glass plates for covers, splinters, cotton wool, soapsuds, a jar of oxygen or apparatus and material for preparing a small quantity of oxygen, wash-bottle containing sulphuric acid, large glass tubes of different diameters and different lengths, small piece of platinum foil, pieces of iron and copper wire.

Experiments with Hydrogen :

1. Slip a glass plate or a piece of paper under the mouth of one of the bottles filled with hydrogen, raise it out of the water and bring it *mouth downwards* over the flame of a spirit lamp.

Where does the gas burn? Why?

Account for the slight explosion when it ignites. What is the color of the flame as it burns up into the bottle? What collects on the inside of the bottle as the gas burns?

2. Remove another bottle from the water as before, and, holding it *mouth downwards*, thrust up into it a burning splinter.

Did the splinter continue to burn? Did the gas burn? Where?

Draw the splinter out slowly. Does it ignite again? Where?

Is the gas a supporter of combustion or a combustible substance as these terms are ordinarily used?

3. Arrange apparatus as in Fig. 13; the delivery tube in this case is a bent glass tube drawn to a fine point; pass the gas coming from the hydrogen generator through sulphuric acid in a wash-bottle for the purpose of removing water vapor, a process often spoken of as

"drying a gas."* After *testing carefully*, to be sure that all the air is expelled, light the gas that issues from the tube.

What is the color of the flame when first ignited, and a moment afterwards?

Bring mouth downwards over the flame a large, cool, dry bottle or beaker. What collects on the inside of the vessel?

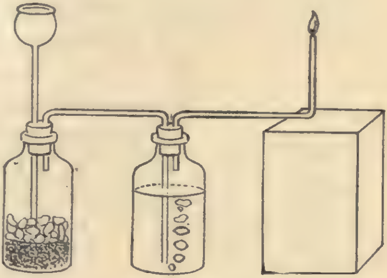

FIG. 13.

4. Bring down over the jet in experiment 3, a number of glass tubes of different lengths, diameters and thicknesses of glass.

What change takes place in the flame? Explain.

What collects on the inside of the tubes? Explain.

Does the size of the flame make any difference in the note heard?

The note is produced by a series of explosions made when hydrogen burns. The pitch depends on the rapidity of these explosions, but the note is modified by the length and diameter of the tube and the thickness of the glass. This is known as the "chemical harmonicum."

5. Roll a piece of platinum foil on a knitting needle, insert it in the end of a glass tube, and seal the glass around the platinum by heating it in the flame of a

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*The drying agents commonly used in the laboratory are sulphuric acid as above, and calcium chloride in a U-shaped tube. The chloride absorbs the moisture, forming a liquid. Such a solid is said to be *deliquescent*. See p. 81.

spirit lamp. Bend the tube, as in Fig. 14. Replace the glass jet by this platinum-tipped one, ignite the gas and note the color of the flame.

FIG. 14.

Hold in this flame a piece of iron wire. Hold a similar piece in the flame of a spirit lamp.

Hold in the flame a piece of copper wire. What color does copper give to the flame?

What do you learn regarding the hydrogen flame?

Would the flame be hotter if the hydrogen were burned in pure oxygen?

Glass contains sodium as one of its constituents, and when it is heated a little of the sodium vaporizes and gives a yellow color to the flame. Platinum is not vaporized nor melted by the heat of the flame, and therefore does not color it, while copper does.

In schools provided with an oxyhydrogen burner (Fig. 15) and gas holders, the intense heat of the oxyhydrogen flame should be demonstrated. A description of the necessary apparatus may be found in Remsen's "Chemistry, Advanced Course," p. 55, and Ostwald's "Principles of Chemistry," p. 103. The

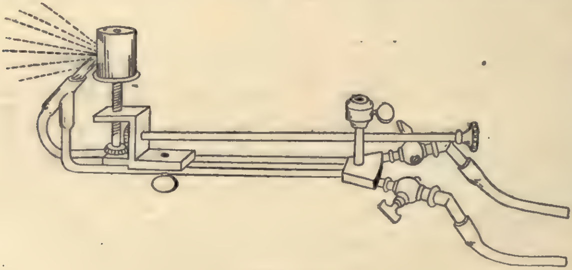

FIG. 15.

experiment with the oxyhydrogen burner is too dangerous to be attempted without proper apparatus.

6. Suspend a bottle containing hydrogen, mouth downwards, or let it stand mouth downwards on a piece of wire gauze for three minutes and test for hydrogen with a burning splinter.

Did the bottle still contain hydrogen?

7. Fasten the delivery tube of the hydrogen generator, by means of a piece of rubber tubing, to a thistle tube, or a clay pipe, the mouth of which is loosely filled with cotton wool.* Use this to make some soap bubbles full of hydrogen.

Is hydrogen lighter or heavier than air? How must it be poured from one vessel to another?

8. Hold a jar vertically in the left hand with its mouth downwards. With the right hand go through the motions of "pouring upwards" from a jar full of hydrogen into the other, Fig. 16. Then test the contents of the latter for hydrogen by a lighted match.

9. Place a bottle containing hydrogen, mouth downwards over the mouth of a bottle filled with air, remove the cover of the hydrogen bottle and allow both to remain this way for about ten minutes. Slip covers between the mouths of the bottles, remove the upper one and drop a lighted match into each.

FIG. 16.

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*The cotton wool must be placed in the mouth of the tube to filter out the fine drops of acid carried over by the hydrogen. These fine drops of acid, touching the soap bubbles, cause them to burst.

When the hydrogen is placed over the air, it gradually mingles with the latter until the whole space contains a homogeneous gaseous solution, although the hydrogen is the lighter gas. This process is called *diffusion*, and this property of diffusing into other gases is common to all substances in the gaseous state.

It is possessed in fact by *all* substances, in varying degrees, when brought in contact with a material which dissolves them. This can be readily seen when water is poured over copper sulphate, in a jar, and the vessel left undisturbed for a few days. The blue color will then be seen to have risen in the jar, and eventually the liquid will become uniform throughout.

10.[*] Fill a soda water bottle over water, two-thirds of its volume with hydrogen and the remaining one-third with oxygen. Bring a lighted match to the mouth of the bottle.

Why should all air be out of the generator before hydrogen is collected?

11. Partly fill a eudiometer over water with hydrogen and oxygen in the proportions by volume of 2:1. Attach an electric cell to an induction coil, connect the wires of the secondary to the platinum wires inserted near the top of the eudiometer.[†]

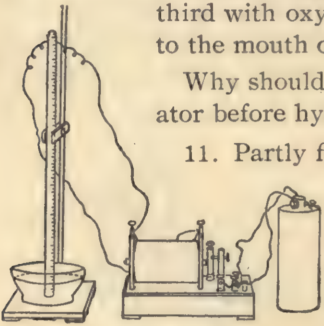

FIG. 17.

Describe and explain what takes place.

In what proportions *by volume* have the hydrogen and oxygen combined?

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[*] The teacher should perform this experiment for the class, because of the danger of explosion.

[†] A few minutes should elapse before exploding, so that the gases may be thoroughly mixed. The difference in the levels of the water inside and outside the eudiometer has not been considered in this experiment.

Repeat the experiment, but take *equal* volumes of hydrogen and oxygen. Measure the gas remaining; test it with a glowing splinter.

What is it? What volume of oxygen has combined with one volume of hydrogen in this experiment?

Repeat again, with *three* volumes of hydrogen to *one* of oxygen. What volume of gas remains? Test it with a lighted match.

What conclusion can be drawn from these three experiments?

By methods described in the text books on Physics, the weights of equal volumes of hydrogen and oxygen measured at the same pressure and temperature have been found to stand in the ratio of 1 to 16. From this information, and the results of the eudiometer experiments, calculate the proportions *by weight* in which hydrogen and oxygen combine. Express the composition thus found in *percentages* of oxygen and hydrogen.

Other Methods of Preparing Hydrogen.

SUPPLIES: Large test tubes, zinc, sulphuric acid, tin foil, potassium, sodium, blue and red litmus paper, a piece of gas pipe, iron turnings, rubber corks, flask, retort stand, glass tubing, delivery tube, some means of heating the iron tube strongly.

Experiments:

1. **Potassium, or Sodium, and Water.** Wrap a piece of potassium or sodium about the size of a small pea in tin foil, drop it in a vessel containing water, and place over it a large test tube also filled with water, as in Fig. 18. When action ceases, cover the mouth of the test tube, raise it out of the water and bring a lighted match to the mouth.

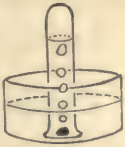

FIG. 18.

2. Drop a small piece of potassium on the surface of water in an evaporating dish.

Note carefully what takes place.

Test the water in the dish with red and blue litmus paper. Is the liquid acid or basic?

Hydrogen is produced so rapidly, when the potassium touches the water, that the force of it throws the potassium a short distance, thus causing it to move about rapidly on the surface. The heat produced is sufficient to raise the temperature to the point at which the hydrogen ignites and the flame is colored by vaporized potassium.

3. Repeat experiment 2, but instead of potassium use sodium.

Can you suggest any reason why the hydrogen in this case does not take fire?

Press a piece of blotting paper just below the surface of the water in the evaporating dish, and into the hollow of this drop a small piece of sodium.

Heat the water in the evaporating dish until it is quite hot, and drop a small piece of sodium on it.

Why does the hydrogen take fire under these conditions? What is the color of the flame?

Test the liquid with red and blue litmus paper.

4. **Steam and Red Hot Iron.** Nearly fill an iron tube (a piece of gas pipe or an old gun barrel) with iron turnings. Connect one end, by means of a cork and glass tube, with the neck of a flask containing water and supported on a retort stand; fit the other end with a

cork and delivery tube which leads to the mouth of a bottle filled with water and inverted in a pneumatic trough, Fig. 19.

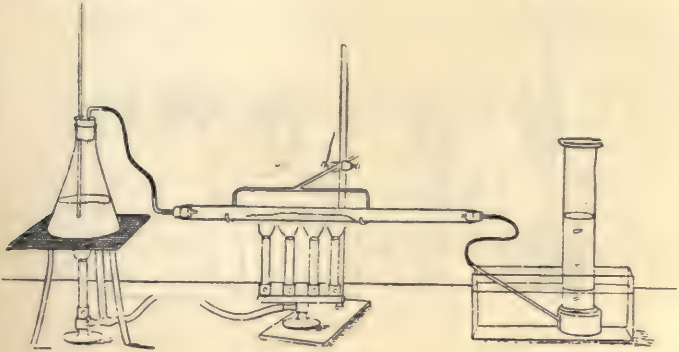

FIG. 19.

Heat the iron tube strongly, boil the water in the flask, collect and test the gas that comes off.

Reduction by Hydrogen.

SUPPLIES: Copper oxide, mercuric oxide, lead oxide, iron oxide, hard glass tubes, connecting tubes and corks, mortar and pestle, hydrogen generator.

Experiments:

1. Place some black copper oxide in a hard glass tube and connect one end, by means of a cork and tube, to a hydrogen generator. Arrange apparatus as in Fig. 20, and pass a current of dry hydrogen through the tube.

After assuring yourself that all the air has been removed from the tube (Why?), heat the oxide in the current of hydrogen.

What escapes at the open end? Hold a cold glass plate near it. What is seen on the sides of the tube?

Allow the hydrogen to pass over the heated oxide for some time, then remove the lamp and allow it to cool down, with the hydrogen still passing over it. When cold, empty the contents into a mortar and rub with the pestle.

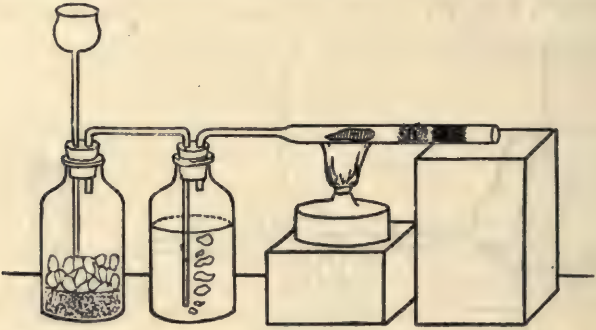

FIG. 20.

What is seen in the mortar?

The copper oxide may be prepared by heating copper in oxygen, and contains both of these substances. When heated in hydrogen, the oxygen unites with the hydrogen to form steam, leaving the copper behind.

When the oxygen combined with the metal to form the oxide, the process was termed *oxidation*, but when the oxygen is removed from the oxide, leaving the metal, the process is called *reduction;* thus, oxygen is an *oxidizing agent*, while hydrogen is a *reducing agent*.

When one hundred grams of copper oxide are heated in hydrogen until all is reduced, the copper remaining weighs 79.9 grams and the water formed weighs 22.6 grams. How much oxygen has united with the hydrogen? How much hydrogen has united with the oxygen?

Compare the proportions of oxygen and hydrogen in water thus found with those calculated at the close of Exp. 11, p. 62.

2. Repeat the experiment, but instead of copper oxide, use mercuric oxide, lead oxide or iron oxide.

The reduction of iron oxide by hydrogen is particularly interesting, because, as shown in Exp. 4, p. 64, iron and steam, which are the products of the reduction, will themselves react to form the oxide and hydrogen again. We have here an important instance of a "reversible" reaction (see p. 22), which has been the subject of much careful study.

NOTES ON HYDROGEN.

Hydrogen was first prepared by Cavendish in 1766, by the action of dilute acids on zinc or iron. It was named by him "inflammable air." It is a colorless, odorless, tasteless gas, very slightly soluble in water: 100 cc. of water dissolves 1.93 cc. of the gas at 0° C. and 760 mm. Its density as compared with air is 0.06926, or, at 0° C. and 760 mm., one litre weighs 0.0902 g. It diffuses rapidly either when in contact with other gases (Exp. 9, p. 6) or when separated from them by a porous partition. It burns in air or oxygen with a very pale bluish flame, forming water vapor. The hydrogen flame is non-luminous, but is remarkable for its intense heat, especially when burning in pure oxygen. Owing to this intense heat, the flame is frequently colored by the vapor of substances in contact with it. When mixed with oxygen and ignited, it explodes violently, although a mixture of oxygen and hydrogen may be kept at ordinary temperature for an indefinite time without union taking place.

This is an example of a reaction which requires starting. At ordinary temperatures, however, the two gases may be made to combine by the action of "platinum black"—a porous form of platinum. The platinum itself, being unchanged by the process, is said to act catalytically. This action is illustrated by one form of automatic gas-lighter, which contains a small quantity of platinum black, and causes ignition by the heat produced locally, by the reaction between the hydrogen in the coal gas and the oxygen in the air.

Ordinarily we say that hydrogen burns in an atmosphere of oxygen, but as this process consists in the mutual union of two gases, with a proper arrangement of apparatus, oxygen may be burned in an atmosphere of hydrogen.

Hydrogen is frequently made use of to remove the oxygen from substances containing it. Such a process is termed "reduction," and hydrogen is known as a *reducing agent.* (See p. 65.)

Hydrogen is hardly ever found in nature, but as one of the constituents of water it occurs in enormous quantities on the earth. It is also contained in all acids and many carbon compounds, such as fats and oils. By means of the spectroscope it has been shown to exist in vast quantities in a free state in the sun, and the same instrument proves its presence in many stars. It may be prepared by two general methods: (1) by the action of dilute acids upon metals, and (2) from water, by the action of a metal. Although many metals besides zinc are acted on by acids with evolution of hydrogen, this metal is generally the most convenient for the laboratory preparation of the gas. When large quantities are required (for inflating balloons, etc.), iron is used on account of its cheapness. Hydrogen prepared by these methods is rarely pure, and frequently a disagreeable odor is noted owing to impurities in the metals. Experiments 1, 2 and 3, pp. 63, 64, illustrate the action of sodium and potassium on water. Other metals, such as iron, which react very slowly with water at ordinary temperatures, readily decompose steam with the formation of hydrogen. (See Exp. 4, p. 64). Highly heated carbon acts in the same way, and when steam, superheated under pressure, is blown through white hot coal, the latter is oxidized and hydrogen set free. Since the oxide of carbon is also a gas we obtain in this case a product containing two gases, carbon monoxide and hydrogen. This product is of commercial importance and is known as "water gas." (See p. 101.)

Like oxygen, hydrogen has been liquefied and solidified.

CHAPTER VII.—WATER

In ordinary language the term water is used very vaguely for a number of different substances. Rain water, well water, sea water, dirty water, are all spoken of as "water" in every day life, although these liquids differ very noticeably in properties.

The difference in the action of soap on well water and on rain water is familiar to all, and even the color and taste varies with the locality. When these various materials are subjected to certain processes, such as filtration, distillation, etc., which will be understood later, substances are obtained which resemble one another more and more closely in properties, and their composition approaches more and more to that of the substance made by the reduction of copper oxide by hydrogen. (See p. 65). Finally, by repeated processes of "purification," substances are obtained which are identical in all respects and are called *pure water*, or by the chemist, who is especially interested in pure substances, simply *water*. The composition of water thus purified, from whatever source it is obtained, whether from sea water, from the explosion of hydrogen in oxygen, or from the action of hydrogen on copper oxide, is found to be exactly the same in every case, and may be expressed conveniently as 88.8% of oxygen and 11.2% of hydrogen.

Experiment:

Distillation of Water. Arrange apparatus as in Fig. 21. Put some water in the flask A and boil for some time, allowing cold water to flow through the condenser B.

Repeat the experiment, using dirty water or water to which some ink has been added.

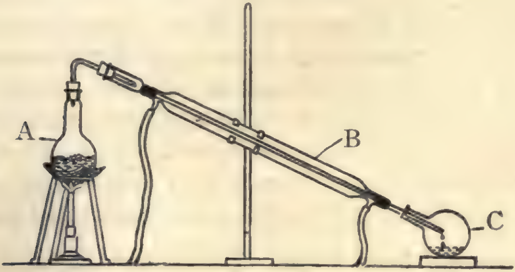

FIG. 21.

What collects in the vessel C in each case?

If hard water is available, its action on soap may be tried before and after distillation.

Water is a type of a very important class of substances, known as *chemical substances* or *chemical individuals*. These terms are applied to materials which are found to have constant composition in spite of all attempts to purify them, or which have the same composition even when prepared by different processes.

Chemical substances may be either *elements* (see p. 35) or *chemical compounds*. Mercuric oxide, common salt, steam, oxygen, magnesium oxide, hydrogen, zinc, copper sulphate or bluestone are also examples of chemical substances which have been met with in the previous chapters. Brine, copper sulphate solution, air, brass, are not chemical individuals because they can be gradually changed in composition by processes of "purification." Brine, for instance, can be gradually purified until on the one hand pure water is obtained and on the other pure salt. Air can be gradually altered in composition and properties until pure oxygen and another chemical individual called nitrogen, which will be described later, are obtained.

Homogeneous substances, like air, brass, brine, etc., which are not mechanical mixtures and which are not chemical individuals, are called *solutions*. Their composition and properties may be varied gradually, *e.g.*, weak and strong brine, various kinds of brass containing varying proportions of zinc and copper. This class of substances will be discussed more fully later.

Properties of Pure Water. Water is a tasteless and almost colorless liquid, although in thick layers, as in clear lakes and rivers, it is seen to have a blue color. Above 4° C. it resembles the great majority of other liquids in expanding with rise of temperature, but below this point it forms an exception to the general rule and expands as the temperature falls.

On account of its cheapness and the ease with which it may be purified, water is used as a standard substance for many purposes. The weight of one cubic centimetre of water at 4° C. has been taken as the unit of weight, the gram; and its freezing point and boiling point define the zero and 100 of the centigrade thermometer. (See p. 24.)

When water is cooled to a certain temperature a reaction takes place, a large quantity of heat is used up, and an entirely different substance called ice is produced. When water is heated, its volume and its other properties gradually change (see pp. 10 and 11) until at a certain point a new substance, steam, is produced, with properties quite different from those of water. These reactions have been discussed on pages 24 and 25.

Water, ice, and steam can thus be converted one into the other by simply heating or cooling. What is "contained" in the one must therefore be contained in the other. This is expressed by saying that these three substances have the same "composition." From this illustration, it is obvious that substances of the same composition may have very different properties. Oxygen and ozone furnish another example, as also the case of graphite, diamond and charcoal, which will be discussed later.

Composition of Water. The following experiment to show the composition of water is given in addition to those already described under hydrogen.

Arrange apparatus as shown in Fig. 22. (A) is the upper part of a large bottle[*] with a small neck, which is closed tightly with a rubber cork having two holes; through these holes are passed short copper wires attached to platinum strips for electrodes, and the holes are plugged tightly with glass stoppers. Nearly fill the vessel (A) with water to which is added one-tenth of its volume of sulphuric acid; fill two graduated tubes or two test tubes of the same size with acidulated water, invert and support one over each electrode. Connect the wires to a battery of four dry cells, in series, and allow the current to pass for some time.

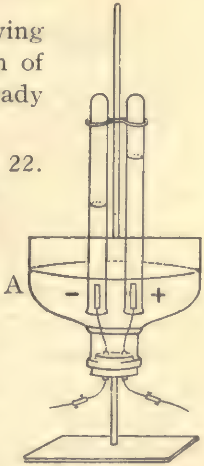

FIG. 22.

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[*] To cut a bottle in two parts, tie a cotton string, saturated with coal oil, around the bottle at the place where it is to be cut, wipe off any oil that may have got on the bottle, ignite the string, and slowly rotate the bottle. When the circle is complete, and the string ceases to burn, either plunge the bottle in cold water, or scratch it with a file.

What is the proportion by volume of the gases in the tubes?

Disconnect the wires from the battery, place the thumb over the mouth of the tube containing the most gas, invert it and thrust into it a lighted splinter. What gas does it contain?

Invert the other tube in the same way, and insert a glowing splinter. What gas is in the tube?

REACTIONS OF WATER WITH OTHER SUBSTANCES.

Water is a compound of hydrogen and oxygen, and in some of its reactions with other substances, hydrogen is a product, in others, oxygen; in by far the greater number, the water simply unites with the other substances to form a new chemical compound or a solution.

Examples of the first type are afforded by the reaction of water with sodium (Exp. 1, p. 63) and by that of steam with red hot iron (Exp. 4, p. 64).

An example of the second type will be met with later in the reaction between chlorine and water in sunlight.

A typical representative of the third class has already been studied in the reaction which bluestone undergoes on heating. As Exp. 9, p. 11, shows, steam is given off and a white powder is left, which unites with water to form bluestone again. There are many salts of this kind, called *hydrated salts*, from which water may be easily obtained by heating, leaving either *anhydrous salts*, or salts containing a smaller proportion of water. These will again combine with water, re-forming the original hydrated salts. (See reversible reactions, p. 22.)

SUPPLIES: Beaker, evaporating dish, Glauber's salt (crystallized sodium sulphate), washing soda (crystallized sodium carbonate), copperas (crystallized ferrous sulphate).

Experiments:

1. Dissolve some alum in a beaker, and divide the solution into two parts. Heat one portion in an evaporating dish over a spirit lamp, until the mass ceases to bubble, and evaporate the other until but little liquid is left. Set both away until next day.

Dry, by pressing between filter paper, a portion of the crystals obtained from the second dish, and then heat them in a test tube.

Try to dissolve each residue (left from the previous day) in cold water. Taste each. Is there any difference? Is there any difference in their appearance?

2. Leave some Glauber's salt, some washing soda, some copperas, exposed to the air for a day.

Describe the changes that have taken place in each.

Many compounds when deposited in the form of crystals from a solution in water, are combined with certain constant quantities of water (therefore chemical individuals). Such crystals are often said to "contain" water, although, since the crystals do not feel moist, it is evident that water is not present in the ordinary sense. It is only because of the ease with which water may be produced from them by heat that they are said to "contain" water at all. (See discussion of word "contain," p. 35.) Owing to the fact that water is a constituent of many crystals, the term "water of crystallization" has frequently been used, although it is by no means essential to the formation of crystals, many substances crystallizing, even from aqueous solutions, in the anhydrous form.

When water is produced from the crystals by heat, the crystals generally crumble to a fine powder, which, although not crystalline to the naked eye, may often be seen by the microscope to consist of minute crystals of another compound with different properties. Alum from which water has been given off, or burnt alum, as it is called, has a different taste and appearance, is but slightly soluble in water, and in fact most of its properties are different from ordinary alum. (See Exp. 1 above.) When copper sulphate is heated, the "water of crystallization" is given off and the color changes. (See Exp. 9, p. 11.)

Some substances from which water has been produced by heating, such as anhydrous copper sulphate, will, if exposed to air, absorb the moisture contained in the atmosphere, and are often used as "drying agents." On the other hand there are many crystals, such as Glauber's salt, washing soda and copperas, which "lose water of crystallization," even at ordinary temperatures, in a dry atmosphere. Such crystals are said to be *efflorescent*, or are said to undergo *efflorescence*

Salts are not the only substances which react with water in the manner described. The reaction of quicklime with water to form slaked lime is quite analogous and may be "reversed" by raising the temperature. Substances like slaked lime, which are formed by the combination of water with the oxide of a metal, are called *hydroxides*.

SOLUTIONS.

A large number of solutions have been met with in the preceding experiments, and some of their properties have been noted. The most common are solutions which "contain" water; but some, like the varnishes, contain alcohol, and on page 8 a solution containing carbon disulphide was prepared.

The solutions form a very important class of substances. The following experiments will illustrate their chief properties.

SUPPLIES: Test tube, beaker, spirit lamp, rubber stopper, glass tubing, thermometer, balance, graduate (burette), caustic soda, caustic potash, iodine, alcohol, carbon disulphide, common salt, ammonium nitrate, copper sulphate, ammonium hydrate.

Experiments:

1. Into a little water in a test tube drop a small piece of caustic potash, or caustic soda. Note any change in temperature; very carefully taste the solution.

2. Shake a small piece of iodine with some alcohol in a test tube. Repeat the experiment, using carbon disulphide* instead of alcohol. Note the color of the iodine and of the products formed in the two cases.

3. Into a test tube measure accurately 40 cc. of water, and into this slowly drop exactly 5 cc. of sulphuric acid. Note the change in temperature. Carefully taste the solution. Measure the volume of the solution.

Repeat the experiment, using 10 cc. of sulphuric acid.

4. Repeat experiment 3, using alcohol instead of sulphuric acid.

What evidence is there that the substances formed in these experiments are not merely mixtures of the original materials? (See p. 12.)

Note. While some of the pupils are performing the following experiments using table salt, others might use ammonium nitrate and others copper sulphate. There

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*** Beware of flames, see note, page 8.**

are two series of experiments: one requiring 10 grams and the other 20 grams of the salt used, for every 100 cc. of water.

5. Weigh a beaker; measure into it 100 cc. of water and weigh again. Weigh out 10 grams of salt; take the temperature of the water in the beaker, put the salt into it, stir with the thermometer and note any change of temperature. Taste the solution. Weigh the beaker and the contents. Measure the volume of the solution.

6. Find the specific gravity of the solution by weighing equal volumes of it, and of water, and comparing the weights

7. Put the solution again in the beaker and placing it on a stand, heat it over the flame of a spirit lamp until it boils. Note the temperature at which it boils freely.

8. Find the temperature at which a similar volume of water boils freely

Note. The boiling point of the solution and the boiling point of water should be found during the same lesson period, as a difference in the atmospheric pressure will cause a difference in the temperature at which water boils. (See p. 24.) The same thermometer must be used, as different thermometers are often not graduated exactly alike. A few scraps of broken clay pipe or unglazed earthenware placed in the beaker will prevent superheating, and the boiling will be more uniform.

9. Repeat experiments 5, 6, 7 and 8, using 20 grams of the salt for every 100 cc. of water

Compare the properties of the solution containing 10 grams of the salt with those of the solution containing 20 grams.

In all the above experiments, it is clear that some reaction has taken place between the substances put together. This was evidenced by the heat given out in the case of the potash; by the obvious difference between the colors of the solutions of iodine and the color of the iodine itself; by the change in the boiling point of the water on addition of the salt; and in the case in which the measurement was made, by the difference between the volume of the solution and the sum of the volumes of the constituents. It could, in fact, be shown that in every case the properties of the substances produced were different from what would be expected were the materials merely "mechanically mixed." Further, the experiments with different proportions of table salt (or ammonium nitrate or copper sulphate) showed that the composition of the substances prepared could be varied at will, and that their properties changed with their composition. The same might have been proved in the other cases also; this is important in the definition of "solution," page 82.

Solutions may also be formed by the action of a gas on a liquid.

10. Partly fill a test tube with water and warm it over the flame of a spirit lamp. Note the small bubbles of air that appear.

11. Into a large test tube put a small quantity of strong ammonia. Close with a perforated rubber stopper, through which runs a long glass tube, Fig. 23. Heat the liquid in the test tube until it boils; invert, and quickly dip the end of the glass tube under water.

FIG. 23.

Explain fully what occurs.

Unsaturated, Saturated and Supersaturated Solutions.

SUPPLIES: Powdered alum, copper sulphate, calcium sulphate (plaster of Paris), test tube or beaker, evaporating dish or watch glass.

Experiment:

Into each of three large test tubes or beakers measure 20 cc. of water. To the first add 2 g. of powdered alum; to the second 2 g. of powdered copper sulphate, and to the third 2 g. of calcium sulphate (plaster of Paris). Shake each and allow them to stand for a few minutes.

What evidence have you that each dissolves? If the third appears doubtful, transfer a small drop of the clear liquid, by means of a glass rod, to a watch glass and evaporate to dryness.

Is the evidence of solution conclusive? Evaporate a drop of the water used.

To the first and second solutions, add alum and copper sulphate respectively, until the salt ceases to dissolve.

The solutions are now said to be *saturated* with the salt. In what condition were they before adding the last portion of the solid?

Heat these solutions and add alum and copper sulphate respectively, until no more will dissolve even when the temperature is nearly at boiling point. Pour the clear solutions into a clean test tube freshly washed out with warm water; plug the mouths with cotton wool and set away where they will not be disturbed until next day.

Examine the contents of the tubes before disturbing them. What has occurred? What does the result of this experiment teach regarding solutions? Does the concentration (see p. 81) of a saturated solution vary with the temperature?

If crystals have not formed in the alum solution, remove the plug and drop in a *very small crystal* of alum. What occurs?

Explanation of Some Terms.

A saturated solution of any substance at any given temperature is a solution that may be brought into contact with that substance without any change occurring. An *unsaturated solution* is one that will dissolve more of the substance when brought in contact with it, while a *supersaturated solution* is one from which the dissolved substance will separate out when placed in contact with some of that substance. (See explosive reactions, page 18.) It is clear, therefore, that a supersaturated solution can exist without change, only when all traces of the substance with which it is supersaturated are excluded from contact with it. A *dilute* solution is one that contains a relatively small proportion of the substance dissolved, while one containing a relatively large proportion is termed a *concentrated* solution. The *concentration* of a solution is the weight of the dissolved substance contained in a unit weight of the solution. Thus a weak solution may be said to have a smaller or lower concentration than a strong one; and one solution may be said to be twice as concentrated as another when a given weight of it contains twice as many grams of the dissolved substance as the same weight of the other.

Deliquescence.

SUPPLIES: Watch glass or glass plate, concentrated sulphuric acid, caustic soda, caustic potash, calcium chloride, zinc chloride.

Experiments:

1. Leave one drop of concentrated sulphuric acid over night on an uncovered watch glass.

2. Leave some caustic potash, caustic soda, calcium chloride, and zinc chloride exposed to the air for some hours.

3. Arrange fresh quantities of the substances used in 1 and 2, and breathe on them for a few minutes.

The materials used in these experiments are said to be deliquescent. They react with water from the atmosphere, forming very concentrated solutions which absorb more and more water until equilibrium is established. On account of this property, deliquescent substances may be used for "drying" gases. (See Exp. 3, p. 58.) Compare the behavior of anhydrous copper sulphate (p. 1), and of quicklime (p. 102).

The reason why common salt becomes moist in damp weather is that it contains, as an impurity, a deliquescent salt of magnesium. Clothes wet with sea water, which contains a considerable quantity of magnesium chloride, are difficult to dry for the same reason.

Solutions and Chemical Individuals.

The solutions so far discussed have been, in almost all cases, liquids; indeed, in common language, the word solution is never used in any other sense. The solids and gases are said to be dissolved in the water, alcohol, etc., and the latter are spoken of as the solvents. When a solution is formed from two liquids it is often spoken of as a mixture, *e.g.*, of sulphuric acid and water. This practice is a bad one, and will not be adopted in this book, as it might be thought that a mechanical mixture was meant.

The word solution, however, is now used in a wider sense, and is applied to all substances (not mechanical mixtures) whose composition may be gradually or continously varied, that is to say, the substances which form them may be combined in gradually or continuously varying proportions to form a continuous series of substances. The properties of these substances vary gradually and continuously like their composition.

By weighing out the proper quantities of salt and water, for instance, a solution (brine) may be prepared containing 1 gram of salt to 100 grams of water, and another containing 2 grams of salt to 100 grams of water. Others may be prepared intermediate in composition between these, for instance, with 1¼, 1½, 1¾ grams of salt to 100 grams of water, and others again intermediate between these last, in infinite number. The properties of all these solutions vary continuously from one end of the series to the other.

Solutions may be gaseous or solid, as well as liquid. Glass is an example of a solid solution ; so also are the various kinds of brass (made by melting together zinc and copper in different proportions) and many other alloys.

Again, if we bring together oxygen and nitrogen, the resulting gas, although often spoken of as a mixture, is in reality a solution, since the properties of the gas produced may be shown to be different from the sum of the properties of the original constituents, and the composition may be varied, giving a series with continuous variation in properties. Thus air is a solution of oxygen and nitrogen.

Chemical Individuals.

A *chemical individual* differs from a solution in that it is not a member of a series of substances as described above. The quantity of oxygen that unites with one gram of magnesium is always the same, even though the temperature and pressure, at which the combustion of the metal in oxygen takes place, be varied. Oxygen and mercury combine in the proportion of one gram of oxygen to 12.5 grams of mercury to form the red powder of Exp. 1, p. 26, and it is not possible to prepare a compound containing more than one gram of oxygen to the 12.5 grams of mercury. Powders containing more than 12.5 grams of mercury to one gram of oxygen can, indeed, be prepared, but on examination they prove to be mechanical mixtures of the red substance with a black one, and all attempts to prepare homogeneous compounds of oxygen and mercury in which the proportions vary continuously have proved unsuccessful. The red and black oxides are therefore not solutions, as defined above, but chemical compounds or chemical individuals.

When the white powder of anhydrous copper sulphate (Exp. 9, p. 11) is mixed with water, a blue solid substance is formed. If less than 0.564 g. of water is added to every gram of the powder, the product is a mechanical mixture of bluestone and the anhydrous salt; if more than that proportion of water is added, the white powder disappears altogether and some of the bluestone dissolves. It is not possible to prepare homogeneous substances containing water and the white powder in gradually varying proportions; hence bluestone is a chemical compound and not a solid solution.

Similarly, one gram of quicklime is completely slaked by 0.321 g. of water. If less water is added, the product is mixed with unaltered quicklime; if more, with a solution.

NOTES ON WATER.

Water exists in vast quantities on the earth as oceans, lakes and rivers; in the air as water vapor, clouds, fogs, mists; in porous soils and rocks, where it is absorbed and retained in the liquid state; in a few rocks as water of crystallization; in all plants and animals, of which it forms a large proportion of their weight; and in the form of ice and snow covering the colder parts of the earth. The action of water on soils and rocks through solution, erosion and alternate expansion and contraction; its service to man through irrigation, transportation and industrial life, are dealt with under physical and commercial geography. Many of its properties—such as expansion and contraction due to variations of temperature; freezing point and boiling point and their dependence on pressure; evaporation and condensation (distillation); influence of pressure on the melting point of ice; maximum density, expansion on freezing; latent heat of fusion of ice; latent heat of evaporation of water; conduction, convection and radiation of heat—are fully treated of in books on physics.

Pure and Impure Water. Water is never found pure in nature. Rain water contains particles of dust, and also gases which it has dissolved out of the air. When the rain reaches the ground it takes up further impurities from decaying organic matter and from the rocks and soil. Some of this water runs along the surface, collecting into streams and rivers, becoming contaminated by the drainage from the surrounding country, by the sewage from towns along its course, or by the decay of animal or vegetable life in it or along its banks. The rest of the rain that has fallen either evaporates again or filters slowly through the soil, and collecting, forms underground streams or lakes, or comes to the surface as springs or wells. As this water percolates through the soil it loses any particles of organic matter, but dissolves minerals or gases. Sometimes the quantity or kind of dissolved substance is sufficient to give the water a marked taste, or medicinal properties, and it is then termed *mineral water*. The varieties of mineral water depend on the nature and quantity of the substances in solution. Water that contains calcium or magnesium salts (derived from the rocks of the neighborhood) forms a precipitate with soap, and is known as "hard" water; in some cases the mineral matter is deposited on boiling (see page 101), when the water is said to be "temporarily" hard.

Ocean Water contains many salts in solution, chief among which are sodium chloride (common salt), magnesium chloride, potassium chloride, magnesium carbonate and calcium carbonate. As the water evaporates from the surface of the ocean these salts are left behind; the vapor condenses and falls as rain on the land, where it dissolves more mineral matter and carries it to the ocean, where it accumulates.

Drinking Water should be freed from suspended particles of either organic or mineral matter. To effect this, the water supply of cities and towns, when taken from rivers or lakes liable to contamination, is often allowed to filter through layers of sand and gravel several feet in thickness. Such a filter will remove bacteria and all suspended matter, but will increase the quantity of mineral matter in solution, and must be cleaned frequently.

It is especially necessary to prevent contamination of drinking water by decomposing animal matter, as the bacteria invariably present in such water are a fruitful source of disease. The mere appearance of a sample of water is no criterion of its purity in this respect, as water which is clear, sparkling, cool, attractive to the eye and pleasant to the taste may be seriously contaminated by these germs, or may be liable to sudden contamination from some unexpected source, while water which has none of these qualities may be quite harmless.

The only really pure water is that which has been freshly distilled. It has a "flat" taste owing to the absence of gases, derived from the atmosphere, which are always to be found in ordinary drinking water.

CHAPTER VIII.—HYDROGEN PEROXIDE

Supplies: Barium dioxide, hydrochloric acid, solutions of potassium iodide, potassium permanganate, logwood, litmus.

Experiments:

1. **Preparation of Hydrogen Peroxide.** Put some barium dioxide in a beaker, add some water to it, stir and add about a quarter of its volume of hydrochloric acid. Filter.

2. To a paper moistened with a solution of potassium iodide, add a drop of the clear liquid from experiment 1.

3. To solutions of potassium permanganate, logwood, and litmus, add some of the liquid containing hydrogen peroxide.

The liquid obtained in experiment 1 is a solution of hydrogen peroxide (or dioxide). Pure hydrogen peroxide is a clear liquid which contains 5.94% of hydrogen to 94.06% of oxygen, as compared with water, which has been shown to contain 11.21% of hydrogen to 88.79% of oxygen.

Strong solutions of hydrogen peroxide decompose readily into water and oxygen, but weaker solutions may be kept for a longer time. Ordinary commercial solutions contain about 3% of the peroxide, and even these decompose slowly, liberating oxygen. Like ozone it is used as an oxidizing agent. Because of its active oxidizing properties, solutions of hydrogen peroxide are extensively used in medicine as a disinfectant.

CHAPTER IX.—CARBON AND ITS COMPOUNDS

Experiments :

1. Make a little soot by holding a cold evaporating dish for one or two minutes in the flame of a candle, or gas jet. Scrape it off and put it in a clean, dry test tube ; cork loosely and hold in the flame of a Bunsen burner, or spirit lamp, until it all burns away.

The soot " burns up " altogether, that is, disappears without leaving any visible residue, but the experiments described in the chapter on the Conservation of Mass make it very unlikely that there is no product at all. We should therefore investigate the air in the tube to see whether it has undergone any change.

2. When the tube is cool again, pour in a little lime water and shake. What do you observe ?

Pour as much lime water into a similar test tube full of air, and shake as before.

What evidence do these experiments afford that a new gas is formed when soot is burned ?

The experiments 3–8 of Chapter II., pages 30–33, showed that substances which burned in air, burned more brightly in oxygen ; this is true of soot, as may readily be shown if a tube full of oxygen is available. The product of the combustion in oxygen may be treated with lime water, and gives the same result as when air was used. This is important, for it shows that the new gas can contain nothing but oxygen and soot, or carbon, as chemists prefer to call it.

Since all attempts to decompose carbon, or to produce it quantitatively by the combination of two or more different substances have failed, it is called an element. (See p. 35.)

The gas prepared above is a gaseous solution of a chemical compound called carbon dioxide, with the excess of air or oxygen used in its preparation. There are many ways of obtaining this carbon dioxide in a state of purity, some of which will be described in the following chapter, and it may be shown to have the same composition, namely 2.66 gram of oxygen to every gram of carbon, even when prepared by very different methods.

The reaction with lime water may be chosen as a "test" for carbon dioxide, and, further, if any substance on heating, or on burning in air or in oxygen forms a gas which gives the lime water reaction, it may be inferred that the substance in question contains carbon. Of course all that is said on page 30, about the caution necessary in using such tests or "reagents," applies in this case with equal force.

Tested in this way a large number of substances are found to contain carbon—the breath for example, wood, meat and all organic matter, alcohol, coal gas, limestone, baking soda, sugar, coal, coke, charcoal, soot, graphite and diamond, etc.

When the three last named substances are burned in oxygen, nothing but carbon dioxide is formed, 2.66 g. of oxygen being used up for every gram of either of them consumed. They therefore contain nothing but carbon; and it is indeed because of this that the common term "carbon" has been given to all of them. Methods have been found for converting these very dissimilar substances into one another, and will be described in the following paragraphs. This is a further illustration of the phenomenon referred to on page 73, in that several quite different substances have the same composition. When it is an element that exhibits this peculiarity, as in the cases of carbon and oxygen, the various substances are often spoken of as *allotropic modifications* of the element in question.

Charcoal is an amorphous (that is, non-crystalline) variety of carbon. It is obtained by heating organic matter in closed vessels, or by partially burning them in air. The process consists essentially in driving off the volatile matter and retaining the carbon.

SUPPLIES: Test tube, spirit lamp, sawdust or clean leaf of scribbling paper, litmus paper, wire gauze, uncolored woollen cloth, mercury, solution of ammonia, charcoal, copper oxide, sulphuric, hydrochloric, or nitric acid, potassium or sodium hydroxide.

Experiments:

1. Place sawdust or paper in a test tube to a depth of about an inch, and heat it for some time in the flame of a spirit lamp, holding the tube in a horizontal position.

What is seen to issue from the mouth of the tube? Test the liquid condensed on the walls of the test tube with litmus paper. What is left in the tube? Why does it not burn?

Place the substance that is left in the tube on a fine wire gauze and heat it for some minutes over the flame of the lamp. What is left?

2. Repeat the above experiment, using some uncolored woollen cloth instead of the sawdust or paper. What difference is noted in the substances that issue from the tube?

3. Fill a large test tube over mercury with ammonia gas, prepared by heating a strong solution of the gas (ordinary liquor ammoniæ); put into the tube a piece of freshly heated charcoal and allow it to stand for some time over the mercury.

What property of charcoal does this illustrate?

4. In a small hole in a piece of charcoal put some copper oxide and heat it strongly in a blowpipe flame.

What is left? What property of carbon is illustrated? Compare the action of the hydrogen on the heated oxide, p. 66.

5. To a piece of charcoal in a test tube add some strong acid, as sulphuric, hydrochloric, or nitric.

Repeat the experiment, but add a strong base, as potassium or sodium hydroxide, instead of the acid. Is there any reaction in either case?

See if a reaction occurs when the materials are heated.

Wood charcoal is made either in a "charcoal pit" or in a large retort or oven. When made in a pit (Fig. 24), the wood is heaped in a hemispherical pile around a central opening and covered with earth and sod, leaving only a few small draught holes near the bottom. It is then ignited at the centre and allowed to burn until the whole pile is on fire. A smoldering combustion takes

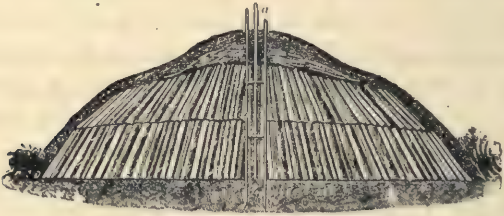

FIG. 24.

place, largely at the expense of the oxygen and hydrogen of the wood fibre, forming water, carbon dioxide and volatile liquids, which escape. The draught holes are then all closed and the pit is carefully covered until the fire smothers and the charcoal is cold. In this method the volatile matter escapes and the yield of charcoal is only 20% of the weight of the wood. It is therefore wasteful and expensive. When made by the dry or destructive distillation of wood in large retorts, or properly constructed ovens, the charcoal amounts to almost 30% of the original wood used, together with nearly 40% of wood alcohol and acetic acid, and 4% of tar. The sale of these by-products covers the expense of the wood used.

To what was the action on litmus paper due in Exp. 1, above?

Animal Charcoal is made by heating bones, or a mixture of blood and sodium carbonate, in a closed vessel. Under the name of *ivory black*, animal charcoal is used as a pigment, especially in making shoeblacking. It is also used to remove the color from sugars, syrups, oils and other liquids colored by organic matter.

6. Partly fill a funnel with animal charcoal (boneblack), and filter through it, two or three times, a solution of dark brown sugar. Evaporate *slowly*, almost to dryness, a little of the clear liquid that comes through the filter.

What property of animal charcoal does this illustrate?

Soot is a comparatively pure form of carbon, formed by the partial combustion, in a limited supply of air, of volatile compounds of carbon and hydrogen.

Coke is made by expelling the volatile matter from soft coal, somewhat as charcoal is made from wood. On a large scale, it is made by heating a special grade of soft coal in huge brick ovens, shaped like beehives, from which the air is excluded after combustion begins. When preparing illuminating gas from soft coal, coke is left in the retorts, the volatile products of distillation being either condensed, or collected by leading them into water, while the gas that passes through may be used to heat the retorts. From the products of the distillation of soft coal a number of valuable materials are obtained, including ammonia, creosote, coal tar, aniline dyes, carbolic acid, benzene, etc. This coal tar is used for making tar paper, roofing material and paving for streets, and as a preservative for wood, metal and stone exposed to the weather.

Coke and charcoal are used in large quantities in smelting ores of metals, such as iron, lead, zinc, copper, etc. They are mixed with the ores, or the products obtained by roasting them, which consist principally of the oxides of the metals; when heated, the carbon reduces the oxides by removing the oxygen. (Compare Exp. 4, p. 92.)

Gas Carbon is deposited on the inside of the retorts used in the manufacture of illuminating gas. It is a black, heavy, hard solid, and is almost pure carbon. It is a good conductor of electricity, and is used extensively in making carbons for electric lights and plates for electric batteries.

Lampblack is made by burning oils, rich in carbon, in a limited supply of air. It is used in making printer's ink and certain black paints. As carbon is acted on, directly, by very few substances (see Exp. 5, p. 92) it is impossible to destroy this coloring matter without destroying the substance on which it has been used.

Graphite is a crystalline form of carbon, found as flaky masses or slabs, particularly in Bavaria, Ceylon, Siberia, New York, and Ontario. Graphite was at one time supposed to contain lead, and is even now incorrectly called "blacklead" and "plumbago." It is used as a lubricant, and in the manufacture of stove polish, protective paints, crucibles and electrodes for electric furnaces. Immense quantities are consumed in making lead pencils. The graphite for this purpose is finely powdered, mixed with more or less clay according to the hardness required, and pressed through perforated plates from which the "lead" issues in tiny rods; these are dried, cut into the required lengths, baked, and inserted in wooden cases.

Graphite is manufactured on a large scale by heating a certain grade of anthracite coal in an electric furnace. Over 800,000 lbs. were manufactured at Niagara Falls in 1902.

Diamond is pure crystallized carbon. When found, diamonds are rough looking stones, and must be cut and polished to give them their peculiar lustre and sparkle. When pure they are colorless, but they are frequently colored owing to the presence of impurities. An impure variety from Brazil, called carbonado, is set in the end of "diamond drills" for boring rocks.

CARBON DIOXIDE.

As already shown, this substance is a colorless gas; it has all the properties common to gases and obeys the gas laws approximately. On the large scale carbon dioxide is prepared by heating limestone, or chalk, which is thereby converted into quicklime and carbon dioxide. This reaction, however, although involving only cheap materials, requires a high temperature, and is not easy to carry out on a small scale. In the laboratory, the gas may be prepared by heating baking soda in a tube like that used for the preparation of oxygen, or by the action of acids on marble ; the latter method is the more convenient.

SUPPLIES : Hydrogen generating apparatus, marble, hydrochloric acid, splinter, candle, lime water, bottles for collecting gas, wash-bottle, glass tubing ; limestone, potassium, sodium or ammonium carbonate ; sulphuric, nitric, or acetic acid ; oyster, clam, or egg shell ; coating from the inside of a kettle.

Preparation of Carbon Dioxide. Into the hydrogen generating apparatus (see Fig. 10, p. 56) put some pieces of marble, cover with water, and pour some hydrochloric acid down the thistle tube. Collect the gas over water, or by the downward displacement of air.

What properties of the gas may be noted ?

Experiments with Carbon Dioxide.

1. Thrust a lighted splinter into a bottle containing the gas.

What two properties of the gas does this experiment show ?

2. Light a candle, lower it into a large bottle or jar and pour down upon it, as though it were water, the gas contained in one of the bottles.

What properties of the gas are shown by this experiment ?

3. Into one of the bottles containing carbon dioxide pour water until it is about one-third full, place the hand

tightly over the mouth and shake vigorously for some time. Invert, still covered, in a trough containing water and remove the hand.

4. Pour a little *diluted* lime water into a bottle containing carbon dioxide, cover with the hand and shake ; still keeping the hand over the mouth shake vigorously for some time.

What precaution does this suggest in connection with the test for carbon dioxide given on page 90 ?

5. Exhale, through a glass tube, into a beaker containing diluted lime water ; continue the process for some minutes.

6. Lower a lighted candle into a bottle containing air, and after it has burned for a short time, remove, add lime water and shake.

7. Repeat the last experiment using a burning splinter.

8. Partially fill a wash-bottle with lime water, and suck ordinary air through it for several minutes.

What constituent in the atmosphere does this show the presence of ? Is there much of it ? (Compare with Exps. 4-7.) Suggest possible causes of its presence in the air. Would you expect the amount of it in the air to be the same in the city and in the country ?

Test the liquid from experiment 3 with litmus. Is the solution of carbon dioxide an acid or a base ? Test some lime water (calcium hydroxide solution) in the same way. Is carbon a metal or a non-metal ? Is it an acid-forming or a base-forming element ? (See p. 36.)

Salts.

In experiment 4, when carbonic acid was added to a solution of calcium hydroxide, a precipitate was formed, called calcium carbonate. This substance belongs to an important class of materials, called salts.

9. To a solution of sodium hydroxide, add dilute hydrochloric acid, until the solution just begins to turn blue litmus red. (By the exercise of a little care a solution may be obtained which will neither turn blue litmus red, nor red litmus blue.) Evaporate a portion of this solution to dryness in a dish.

In this experiment a typical salt (common salt) has been obtained from an acid and a base. Since salt was left behind when the water was evaporated, the liquid is said to be a solution of the salt. On account of the ease with which a salt is generally obtained from its solution, the product of the reaction of an acid on a base is often spoken of as a salt, although in many cases it is a solution that is produced.

Carbonic acid with other bases forms other salts, all of which are known as carbonates, and solutions of calcium hydroxide with various acids give salts known as salts of calcium.

Salts may also be prepared in other ways, as will be shown later.

Calcium Carbonate.

The most important carbonate is calcium carbonate, which occurs in vast quantities as limestone, marble, chalk, calcite and in the shells of many small animals that live in water.

10. Into a test tube put a small piece of limestone. Add hydrochloric acid, and hold the mouth of the tube above the mouth of another test tube containing lime water, as shown in Fig. 25. Shake the tube containing the lime water.

11. Repeat Exp. 10, using any carbonate, as potassium, sodium, or ammonium carbonate, with sulphuric, nitric, or acetic acid.

FIG. 25.

12. Repeat the experiment, placing in the first test tube some pieces of oyster, clam, snail, or egg shell.

13. Repeat the experiment, using pieces of the coating from the inside of a tea kettle.

Mortar is formed from slaked lime, sand and water. After being placed on the wall, it slowly "sets," the process consisting partly in the evaporation of the water into the air, and partly in the combination of the carbon dioxide of the air with the calcium hydroxide (slaked lime), forming calcium carbonate. That this reaction takes place may be shown by placing a piece of old mortar in acid, and testing the gas given off with lime water.

Soda Water.

Soda water is the trade name for a concentrated solution of carbon dioxide in water. It contains no soda.

SUPPLIES: Soda water, lime water, litmus paper, spirit lamp, sand, pieces of clay pipe or unglazed earthenware.

Experiments:

1. Test a little diluted soda water with blue litmus paper and with lime water.

2. Heat some soda water in a beaker and note what occurs. Test the gas given off, by holding a glass rod moistened with lime water at the mouth of the beaker. Boil for about five minutes and again test, as before, with the rod and lime water. When cold, test the remaining liquid with lime water* and with litmus.

3. Pour some soda water into a beaker and leave it without shaking until the bubbling has ceased. Then add a pinch of sand, a piece of a clay pipe or unglazed earthenware.

These experiments show several important properties which are common to all gases. In the first place, when the soda water bottle was opened, bubbles began to rise in the liquid. The space above the liquid in the bottle contained carbon dioxide under a considerable pressure, and when this pressure was reduced, the amount of carbon dioxide dissolved in the solution (or its concentration, see p. 81) decreased. It is a general law that the solubility of a gas in a liquid decreases with decrease in pressure. This furnishes another example of a reversible reaction, for the soda water was originally prepared by bringing water into contact with the gas under high pressure.

The evolution of bubbles in the second experiment shows that the solubility of the gas decreases with increase of temperature. This also is true of all gases.

The third experiment illustrates the behavior of a supersaturated solution of a gas. The action of the sand, or clay, is analogous to that of the crystal in Exp. 1, p. 19.

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* Some samples of soda water contain traces of alkaline carbonates. In such cases there will always be a slight precipitate with lime water, even after boiling.

NOTES ON CARBON DIOXIDE AND CARBONATES.

Carbon dioxide is a colorless gas having a slight taste and odor, and, at the ordinary temperature of the laboratory, is soluble in about its own volume of water. One litre at 0° C. and 760 mm. weighs 1.964 g. It is about one and a half times heavier than air, and on this account may be collected by the downward displacement of air, or may be poured from one vessel to another like water. For the same reason it frequently collects in deep wells and in valleys near volcanoes or lime kilns.

Carbon dioxide is present in the air in variable quantities according to conditions. In places where there is no special source of the gas present, the amount is about 0.05 per cent. of the volume of the air. The amount is increased by animal respiration, by combustion of all substances containing carbon, by decomposition of animal matter, and by volcanic action. This carbon dioxide dissolves in rain and is carried to the earth, where the solution percolates through the soil and dissolves many substances, particularly calcium carbonate (limestone) (see Exps. 4 and 5, p. 97), for although limestone is not soluble in pure water, it is soluble in water containing carbon dioxide, and is therefore found in spring, well, or lake water, where the neighboring country contains limestone rocks. When this water is heated, carbon dioxide is driven off and calcium carbonate is deposited. It is for this reason that tea kettles and boilers of steam engines become encrusted with this substance.

When calcium carbonate is heated, carbon dioxide is driven off and the oxide of calcium (quicklime) remains; this is the reaction which takes place in a lime kiln. When water is added to the quicklime, the two combine, giving calcium hydroxide or slaked lime. (See water, p, 84.)

Experiment:

Place some fresh quicklime in a saucer and pour on it some lukewarm water. What evidence is there that a reaction has taken place?

When slaked lime is added to water, part of it dissolves, forming a solution known as lime water. When carbon dioxide is passed into lime water, the liquid becomes whitish in color owing to the formation of small particles of calcium carbonate.

The change of limestone into quicklime and carbon dioxide is another example of a reversible reaction, analogous to the evaporation of water. When calcium carbonate is heated, calcium oxide is formed and carbon dioxide gas is given off; and just as the pressure of steam from water increases with the temperature, so also the pressure at which carbon dioxide is given off by the calcium carbonate rises as the temperature is increased. If the temperature be lowered again, the calcium oxide and the carbon dioxide unite to form the carbonate, and the pressure is thus decreased. For every temperature there is a corresponding pressure at which the three materials can exist together in equilibrium.

In the process of breadmaking carbon dioxide is liberated by the fermentation produced by yeast, and causes the "rising" of the sponge or dough. The gas may be generated also by the use of baking soda and some acid, such as cream of tartar. The various baking powders are mechanical mixtures of sodium bicarbonate (or some other carbonate) and some substance which forms an acid solution with water. That they are mixtures and not solid solutions, or chemical compounds, is clear from the fact that the constituents when brought together show no evidence of reaction; moreover, the properties of the components are the same as in the unmixed state. When, however, the mixture is moistened, a reaction takes place, and carbon dioxide is given off.

Liquid carbon dioxide is now made in large quantities, by forcing the gas into steel cylinders by means of powerful pumps. It is used to prepare soda water, to extinguish fires, to improve wines, and to produce very low temperatures. When the tap of a cylinder of carbon dioxide is opened the liquid evaporates rapidly, using up heat, and under suitable conditions a portion may be obtained frozen to a snow-white solid.

CARBON MONOXIDE.

When carbon was burned with a free supply of oxygen, a compound called carbon dioxide was produced. If the combustion is carried out under different conditions, with a limited supply of oxygen, an entirely different substance is formed which is called carbon monoxide, for a reason that will be presently explained. This gas is seen burning with a blue flame at the surface of hard coal fires.

SUPPLIES: Oxalic acid, sulphuric acid, flask, retort stand, wash bottle containing a solution of caustic potash or caustic soda, connecting tubes and delivery tube, bottles to collect the gas over water, splinters, lime water.

Preparation. Arrange apparatus as shown in Fig. 26. Into the flask put 10 g. of oxalic acid and 50 cc. of sulphuric acid. The wash-bottle contains a strong solution of sodium hydroxide. Heat the flask gently and collect several bottles of the gas over water.

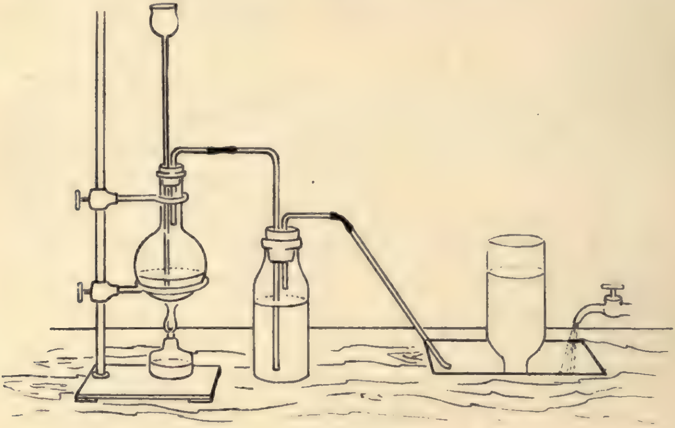

FIG. 26.

What properties of the gas may be noted?

By the action of sulphuric acid on oxalic acid, a gas is given off which contains both carbon dioxide and carbon monoxide. On passing this gas through the solution of sodium hydroxide, the carbon dioxide reacts with the base forming a solution of sodium carbonate (see Salts, page 98), while the monoxide, which does not form an acid with water, passes on through the solution.

Experiments with Carbon Monoxide:

1. Thrust a lighted splinter into one of the bottles filled with the gas.

After action has ceased, again thrust in a lighted splinter. Pour lime water into the bottle and shake for some time.

What is the color of the flame? What is formed when carbon monoxide burns?

2. Fill a bottle with carbon monoxide and oxygen, in the proportions of about 2:1 by volume, and apply a lighted splinter to the mouth.

NOTES ON CARBON MONOXIDE.

Carbon monoxide, or carbonic oxide, as it is sometimes called, is a gas without color, odor, or taste, and is but slightly soluble in water. One litre at 0° C. and 760 mm. weighs 1.250 g. It burns with a bluish flame, forming carbon dioxide. It contains only carbon and oxygen, differing from carbon dioxide in having 1.33 grams of oxygen to every gram of carbon, whereas the dioxide, as stated on page 90, contains exactly twice this proportion. It is for this reason that the terms monoxide and dioxide are employed.

It is formed when carbon is burned in a restricted supply of air. It is also formed when carbon dioxide is passed through highly heated coal. This may be seen in any coal fire. The air enters at the lower grate and the oxygen combines with the carbon, forming carbon dioxide, but this gas, in passing through the upper layers of glowing coal, is reduced to carbon monoxide. At the top of the glowing layer of coal, where the air has free access, the carbon monoxide burns with a flickering bluish flame, forming carbon dioxide.

Carbon monoxide is extremely poisonous, and is particularly dangerous because it is odorless and therefore difficult to detect. Cases of carbon monoxide poisoning frequently occur where coal is burned in a stove with insufficient draught, resulting in the escape of carbon monoxide into the room. Coal gas contains sufficient carbon monoxide to produce death by poisoning.

Carbon monoxide is one of the constituents of coal gas formed by the distillation of soft coal, and is also contained in *water gas* (see p. 111), which is formed by blowing steam, superheated under pressure, through white hot coal. In this reaction the carbon reduces the water to hydrogen and is itself oxidized to carbon monoxide. The carbon monoxide and hydrogen thus produced, burn with a non-luminous flame and may be used for heating purposes. The flame may be rendered luminous by the admixture of vapor from oils, and the gas may be then used as an illuminating gas (see p. 110), but it is objected to on account of its poisonous properties.

COMPOUNDS OF CARBON AND HYDROGEN.

Methane.

SUPPLIES: Dry sodium acetate (made by heating sodium acetate in an evaporating dish until the water of crystallization is driven off), sodium hydroxide (solid), quicklime, large glass test tube (or copper retort), retort stand, delivery tube, bottles for collecting gas over water.

Preparation. Arrange apparatus as in Fig. 27. Into the large glass test tube, or copper retort, place about equal weights of dehydrated sodium acetate, dry sodium hydroxide and powdered quicklime. Fit with cork and delivery tube, heat the mixture and collect the gas in test tubes, or bottles, over water.

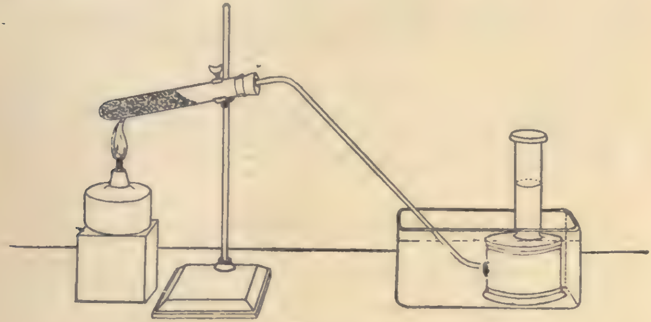

FIG. 27.

In this experiment the test tube should be about three-quarters filled with the mixture, but only the lower portions of the tube should be heated. In this manner the water which condenses on the sides of the glass and runs back, is caught by the powder and absorbed before it comes in contact with the hot glass. The tube should be supported in a nearly horizontal position during the heating.

Experiments with Methane :

1. Ignite one of the test tubes of gas and hold over the flame a cool, dry beaker. What collects on the cool glass ?

After combustion ceases, pour lime water into the test tube and shake.

2. After a couple of test tubes full of the gas have been collected and found to burn quietly, a fine pointed glass jet may be attached to the end of the delivery tube and the gas ignited. Hold mouth downwards over the burning jet a small, dry beaker.

What is formed in the beaker? Does methane burn with a luminous flame? What two constituents of methane are indicated by these experiments?

3. If enough of the gas can be prepared, fill a soda water bottle with one part of the gas and two parts of oxygen by volume. Ignite the mixture.

Notes on Methane.

Methane has no color, taste or odor; it is slightly soluble in water and burns with a pale luminous flame. It is a compound containing only hydrogen and carbon, in the proportion of 0.337 gram of hydrogen to 1 gram of carbon. One litre of the gas at 0° C. and 760 mm. weighs 0.7161 g. A mixture of methane and oxygen explodes violently when ignited, the products of the combustion being carbon dioxide and water.

Methane is a component of the gases evolved from decaying vegetation at the bottom of stagnant water. From this it derives its name, *marsh gas*. It is also frequently found in coal mines, where, on account of its explosive property when mixed with air, it is known as *fire damp*. It generally occurs shut up in cavities and escapes when these are opened in mining. The mixture of water and carbon dioxide formed by the explosion is called by the miners *choke damp*. To prevent such explosions, Davy's safety lamp is used by the coal miners.

There are a great many other compounds of hydrogen and carbon, among which may be mentioned acetylene, benzene, naphthalene. Coal oil, gasoline, paraffin wax, etc., are solutions containing various compounds of carbon and hydrogen.

Acetylene.

Acetylene deserves special mention, as it has recently become of great importance as an illuminating gas.

Experiments :

1. **Preparation.** Fill a bottle with water and invert it in a trough containing water. Roll a small quantity of calcium carbide in tinfoil and insert it under the mouth of the bottle.

What properties of the gas may be noted ?

Remove the bottle and thrust a lighted splinter into it.

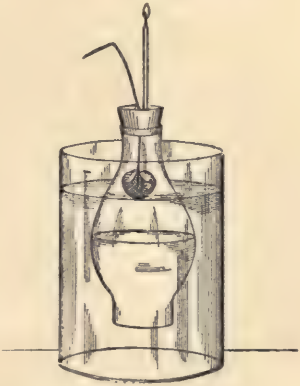

FIG. 28.

2. Arrange apparatus as shown in Fig. 28. A lamp glass is fitted with a perforated rubber cork. The calcium carbide is placed in a small cage made of copper wire gauze, to which is attached a piece of copper wire. The wire is passed through the hole in the cork and held in place by a glass tube drawn out to a fine point. The whole is placed in a large glass vessel or trough of water, as in the diagram. When the air has been driven out (Why ?), ignite the acetylene gas at the top of the glass tube. Hold a cool, dry beaker over the flame.

Notes on Acetylene.

Acetylene is a colorless gas containing 0.0842 g. of hydrogen to every gram of carbon and is lighter than air. It is poisonous if breathed in large quantities, but is much less dangerous than gases containing carbon monoxide. It burns in the air with a smoky flame, that can be made white and very luminous by the use of special burners (Fig. 29), which allow a thin jet of gas to burn with easy access of air. The great luminosity of the acetylene flame is due to its high temperature.


FIG. 29.

Under 40 atmospheres pressure acetlyene may be liquefied at 20° C., but in this condition it may decompose into carbon and hydrogen with liberation of heat, which raises the temperature and causes more rapid decomposition until a violent explosion results.

Acetylene is prepared by the action of water on calcium carbide. The latter substance is prepared by the action of coke on lime, at the very high temperature of the electric furnace. The pure carbide contains 1.66 g. of calcium to 1 g. of carbon, but the commercial carbide is only about 80 per cent. pure.

ILLUMINATING GASES.

Besides acetylene, there are two other gases—artificially prepared, in ordinary use for illuminating purposes—*coal gas* and *water gas*. For the preparation of the latter see page 106.

Water Gas is never used alone for illuminating purposes, but is either mixed with from 60 to 70 per cent. of coal gas, or is "carburetted" by passing it through chambers in which petroleum residues are heated to a high temperature. By this process gases are produced which together with the water gas burn with a luminous flame.

Coal Gas, which is widely used for heating and lighting, is obtained from the distillation of soft coal in large iron retorts. It is a gaseous solution of gases, the composition of which varies with the coal used, the temperature of distillation, and the extent to which it is purified at the works.

The average composition may be given as follows :—

Hydrogen	49	parts by	volume
Methane	34.5	"	"
Olefiant gas and other hydrocarbons	5	"	"
Carbon monoxide	7.2	"	"
Carbon dioxide	1.1	"	"
Nitrogen	3.2	"	"
Total	100	"	"

When the coal gas comes from the retort it contains many other substances, among them being tar, ammonium compounds, hydrogen sulphide, and a larger quantity of carbon dioxide. The coal tar and other volatile liquids are condensed on cooling ; the sulphur is removed by passing the gas over trays of oxide of iron, and the ammonia by washing* with water. The carbon dioxide, not being so soluble in water, is only partially removed.

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* By "washing" a gas is meant bubbling it through some liquid which will remove an impurity by forming a solution with it.

Natural Gas is a combustible gas obtained in many localities from wells bored in the earth. It is used for lighting and as a fuel for heating houses and generating steam in factories. The principal constituent is methane, although it contains other compounds of hydrogen and carbon, together with carbon monoxide.

Luminous and Non-Luminous Flames.

Experiments :

1.* Soak small pieces of cotton wool with alcohol, carbon disulphide, coal oil and turpentine. Place these on plates and ignite.

Note the difference in the flames. Hold a cold piece of porcelain in each.

2. With a bent glass tube drawn out to a fine point, blow air into the side of the flame from the coal oil.

Note the appearance of the part of the flame deflected by the current of air.

3. Compare the appearance of the flame of a Bunsen burner with that of an ordinary gas jet. Remove the lava tip of the gas burner and deflect part of the flame with the current of air as above.

Some substances burn with a luminous flame, others with a non-luminous. In both cases the combustion of the gas or vapor is incomplete in the inner part of the flame, owing to the limited supply of oxygen, and the substances are burnt completely, only when the outer edge of the flame is reached. In non-luminous flames, however, the combustion inside the flame is more thorough than in the case of luminous flames. The combustion inside the flame may be made more nearly

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* This experiment should be performed by the teacher, as these liquids are very inflammable.

complete, either by blowing air into the side of the flame, as in Exp. 2, or by mixing air with the gas before it reaches the jet, as in the Bunsen burner. In this way a luminous flame may be made non-luminous.

What advantages has a non-luminous over a luminous flame for heating purposes?

Other Compounds of Carbon.

In addition to the compounds of carbon with oxygen and hydrogen, there are many thousands of other compounds of carbon containing one or more of the elements oxygen, hydrogen, nitrogen, chlorine, etc. The substances found in plant and animal life consist largely of such compounds of carbon. Starch, cellulose, fats, albumin, sugars, as well as many of their products, such as alcohol, acetic acid and oxalic acid, etc., illustrate the great importance and variety of carbon compounds. Owing to the vast number of these compounds of carbon, and the part they play in the organic world, their study has been made a special branch of chemistry, known as organic chemistry.

CHAPTER X.—A REVIEW OF MECHANICAL MIXTURES, CHEMICAL INDIVIDUALS, AND SOLUTIONS

BEFORE proceeding to the next chapter, in which an extremely important law governing the reactions between chemical individuals will be discussed, it is essential that the classification of substances into mechanical mixtures, chemical individuals and solutions be thoroughly understood. The present chapter contains a review of the characteristics of these classes of substances, and of the distinctions between them.

MECHANICAL MIXTURES.

(Refer also to pages 12 and 15.)

On page 12 it was explained that when two or more substances are brought together without any evidence of reaction taking place, the resulting substance is called a mechanical mixture, or simply a mixture. It is characteristic of a mixture that all its properties are exactly what might be calculated on the assumption that the properties of the components were unchanged by the mixing. (Such meaningless or faulty definitions of mechanical mixtures as "substances that can be separated by mechanical means," or "substances the constituents of which can be dissolved out by certain solvents," etc., should be avoided.)

In many cases, as in that of a mixture of coarsely ground salt with sand, mere inspection is sufficient to enable us to decide that a mechanical mixture is before us. In others, quantitative experiments are necessary; and since it has been found that in general one property of a substance seldom changes without the other properties undergoing a change at the same time, the experimental proof that *one* property of the substance has remained unchanged is often sufficient to show that a mechanical mixture has been formed.

Gunpowder is a substance often taken as an illustration. To decide whether the nitre it contains is mixed or combined with the other ingredients, it is only necessary to see whether some suitable* property of the nitre is altered by the addition of charcoal and sulphur. The solubility of nitre in water, for example, or the amount that will dissolve in 100 grams of water when an excess of nitre is shaken up with water, is found to be 31.2 grams at 20° C. The fact that shaking up an excess of gunpowder with the same amount of water gives a solution containing exactly the same amount of nitre, shows that the solubility of the nitre has remained unchanged, and that, therefore, nitre is present as a constituent of a mechanical mixture. (It does *not* show that *none* of the nitre has combined with (say) the sulphur, nor does it show that the sulphur has not combined with the charcoal. Can you suggest an experiment that would settle these questions?)

In the experiment, p. 13, the fact that the boiling point of the water was unchanged when chalk, etc., were added, showed that water was present *mixed* with other constituents. Did it show that *none* of the water had combined with the chalk?

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* The weight, for example, would obviously be unsuitable. Why?

Experiment:

Repeat the experiment, p. 13, using a little Portland cement instead of the chalk. Does water combine with the cement? Is the water present mechanically mixed with something else? With what?

CHEMICAL INDIVIDUALS AND SOLUTIONS.

(Refer also to pages 76–84.)

As explained on pp. 71 and 72, an investigation of homogeneous substances leads to a very important classification of these into *chemical individuals* and *solutions*. Substances whose composition and properties cannot be altered gradually and continuously are called chemical individuals or (when not elements) *chemical compounds*, while homogeneous substances which are members of series with gradually and continuously changing composition and properties are called *solutions*. Both solutions and chemical compounds differ from mechanical mixtures in that their properties are not those of the substances which they are said to "contain."

The classification of substances into chemical individuals and solutions, like all classifications of natural objects, is in some cases difficult to apply. A full discussion of these difficulties would go beyond the limits of an elementary text book.

Muddy sea water well illustrates the various classes referred to above. In the first place, it is a *mechanical mixture* of solid particles of sand, clay, etc., with a homogeneous liquid. The solid particles may be chemical individuals or solid solutions. The liquid part may be varied gradually in composition, for instance, by evaporation, or by the addition of water, and is therefore not a chemical compound. It is said to "contain" a chemical compound, water, because pure water can be obtained from it, as also other chemical compounds, such as common salt and magnesium sulphate. These substances can be obtained from it by evaporation, and their composition is found to be independent of the composition of the liquid from which they are crystallized. That the brine is a solution of these compounds, and not a mixture, is proved by the fact that the properties are not those of the pure constituents (for instance, the pressure with which brine gives off steam, or its *vapor tension*, at a given temperature, is less than that of pure water at the same temperature). Moreover, when these salts are brought together with water, a change of temperature takes place, showing that a reaction occurs.*

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* That the decision as to whether a substance is a chemical individual or a solution is not such a simple matter as at first sight might appear, is illustrated by the following examples. It has been stated that water is a chemical individual because its composition cannot be altered gradually, but it is evident that by dissolving oxygen in water the proportion of oxygen may be gradually increased, while by dissolving hydrogen in it, the proportion of this substance may be increased. In this case water appears as a member of a series with gradually varying composition. It is only by a careful investigation of the *properties* that it is found that these do not vary continuously.

In the case of a gas a different principle is employed to decide whether it is to be classed as a chemical individual or as a solution.

Although it is possible, for example, to vary gradually the composition of steam, by adding, on the one hand, hydrogen and on the other hand, oxygen, the fact that the properties of steam are different from those of a gaseous solution of hydrogen and oxygen in the proportions in which these substances occur in water, shows that steam is not a solution of these two gases. Further, the fact that there are no two gases known which, on forming a solution might be expected to give a gas with the properties possessed by steam, is a reason for calling steam a chemical compound and not a solution. The question being more easily settled for a solid or liquid, it is often taken for granted that a liquid or solid chemical compound, when changed to the gaseous state, will remain a chemical compound, although the case of ammonium chloride shows that this is not always true.

CHAPTER XI.—LAW OF REACTING WEIGHTS

WITH the help of facts recorded in the earlier part of the book, it will now be possible to illustrate one of the most important and striking generalizations of chemistry, namely, the Law of Reacting Weights. We shall first express this law in words and then show that it is true in a number of examples.

Statement of Law. It has been found as a result of a great number of experiments, that in all reactions involving only chemical individuals, the proportions by weight in which they take part are always simple multiples (such as 1, 2, 3, 4, ½, ⅔, etc.) of certain weights which are constant and characteristic for each chemical substance and are called the reacting weights.

In order to illustrate the law, it is necessary to have the reacting weights of a number of substances. How these are obtained is explained by the following examples :

The formation of water from oxygen and hydrogen is a reaction involving only chemical individuals, and the proportions by weight in which they take part have been found by accurate experiments to be 88.8 g. oxygen and 11.2 g. hydrogen giving 100 g. water (see pp. 62 and 63).

As all three substances are chemical individuals (*i.e.*, are of fixed composition), the *ratios* between the above numbers are independent of the scale on which the experiment is carried out, and the result might be stated in many other ways; for instance:

	1 g.	oxygen	and	0.126 g.	hydrogen give	1.126 g.	water;
or	100 g.	"	"	12.6 g.	" "	112.6 g.	water;
or	16 g.	"	"	2.02 g.	" "	18.02 g.	water,

etc., etc.

As a starting point the reacting weight of oxygen has been arbitrarily selected as 16 g. We find, therefore, that in a reaction involving 16 g., or one reacting weight of oxygen, 2.02 grams of hydrogen and 18.02 grams of water have taken part. Consequently we might select 2.02 grams, 3.03 grams, or 1.01 grams, etc., as the reacting weight of hydrogen, since the weight taking part, 2.02, is a "simple multiple" of any one of these three numbers. (The reasons that determine the selection of one of these multiples in preference to the others will be given later.) Similarly we might take 18.02 grams, 9.01 grams, 36.04 grams, etc., as the reacting weight of water.

By quantitative investigation of other reactions involving only chemical individuals, we can obtain numbers from which to select the reacting weights of other substances. The following table gives the weights of a number of substances which take part in reactions with 16 grams of oxygen. They are all obtained, as shown above in the case of water, from the results of actual experiments.

16 g.	oxygen	and	2.02 g.	hydrogen give	18.02 g.	water.
16 g.	"	"	18.02 g.	water from	34.02 g.	hydrogen peroxide.
16 g.	"	"	6 g.	carbon give	22 g.	carbon dioxide.
16 g.	"	"	12 g.	carbon give	28 g.	carbon monoxide.
16 g.	"	"	4.01 g.	methane give	11 g.	carbon dioxide and 9.01 g. water.
16 g.	"	"	40.1 g.	calcium give	56.1 g.	calcium oxide.

From this table we see that in one case 6 grams and in another 12 grams of carbon have taken part in a reaction with 16 grams of oxygen. Consequently we may take 6 grams or 12 grams or any simple multiple of them as the reacting weight for carbon. Similarly we may take for :—

			Particular multiple selected as reacting weight.
Hydrogen,	2.02	g., or any simple multiple.	1.01 g.
Water,	18.02, or 9.01	" " "	18.02 "
Hydrogen peroxide,	34.02	" " "	34.02 "
Carbon,	12, or 6	" " "	12 "
Carbon monoxide,	28	" " "	28 "
Methane,	4.01	" " "	16.04 "
Carbon dioxide,	11, or 22	" " "	44 "
Calcium,	40.1	" " "	40.1 "
Calcium oxide,	56.1	" " "	56.1 "

The last column gives the particular multiple of the experimental weight that, for reasons to be given presently, has been selected as the reacting weight for each substance.

The table itself shows at a glance two examples of the Law of Reacting Weights, for the proportions in which carbon takes part in the two different reactions involving 16 g. of oxygen, namely, 12 g. and 6 g., are *simple multiples* of its reacting weight, *i.e.*, 1 × 12 g. and ½ × 12 g.. So also, in the case of carbon dioxide, 11 g. and 22 g. are both simple multiples of its reacting weight, *i.e.*, ¼ × 44 g. and ½ × 44 g.

Furthermore, if the "Law" is true in all cases, the reacting weights given in the above table, or some simple multiple of them, will give the proportions in which these substances take part in all other reactions *which involve chemical individuals only*. The following examples, so far as they go, show that this is the case, and in all the thousands of chemical reactions that have been studied quantitatively, this law has been found to hold.

(1) On page 84 it was stated that 1 g. calcium oxide reacts with 0.321 g. water, giving 1.321 g. calcium hydroxide. If we multiply these numbers by 56.1 (to bring them to the basis of one reacting weight of calcium oxide), 56.1 g. calcium oxide and 18.02 g. water give 74.12 g. calcium hydroxide.

It will be noticed that in this reaction the calcium oxide and water take part in the proportion of their reacting weights given in the table. This reaction shows further, that 74.12 g. must be a simple multiple of the reacting weight of calcium hydroxide.

(2) The composition of marsh gas was given on p. 108 as 1 g. of carbon to 0.337 g. hydrogen. This indicates that marsh gas might be made indirectly by the reaction of 1 g. carbon and 0.337 g. hydrogen, giving 1.337 g. marsh gas. Or, multiplying through by 12 (to bring these numbers to the basis of one reacting weight of carbon), 12 g. carbon and 4.04 g. hydrogen give 16.04 g. marsh gas.

In this case it is seen that 4.04 g. is a simple multiple of the reacting weight of hydrogen (4 × 1.01), and the 16.04 g. for marsh gas is a simple multiple of its reacting weight (1 × 16.04).

(3) Example 1 showed that the reacting weight of calcium hydroxide must be 74.12 g., or some simple multiple of this weight. Now this substance combines with carbon dioxide in the proportion of 74.12 g. to 44 g., that is, in the proportion of *one* reacting weight of calcium hydroxide to *one* reacting weight of carbon dioxide.

APPLICATION OF THE LAW OF REACTING WEIGHTS.

The law of reacting weights very much simplifies the task of remembering the composition of chemical substances, and the proportions in which these react with each other. For if the reacting weights of various substances are once determined and recorded in tables, it is only necessary to remember the *multiples* of them that take part in each reaction, in order to know the proportions in which the substances take part in the reaction in question ; and as these multiples are generally small round numbers, or very simple fractions, the tax on the memory is not very great.

Moreover, a system of symbols has been invented which is a still further aid to the memory. The reacting weight of each element is represented by a simple symbol, generally one or two letters suggested by the name of the element. For instance—

16 grams oxygen is represented by O
1.01 " hydrogen is " " H
40.1 " calcium is " " Ca
12 " carbon is " " C

These symbols may be used further to express the reacting weights and the quantitative composition of *chemical compounds*. The reacting weight of water, for instance, as given in the table, is 18.02 grams, and this weight of water contains 16 grams of oxygen (represented by O) and 2.02 grams of hydrogen (twice the amount represented by H). These facts may be indicated by writing 2H and O together, thus HHO or H_2O. The symbol so formed, which is called a " chemical formula," stands for 18.02 grams of water. It tells us that 18.02 is the reacting weight of water, *i.e.*, that (in reactions involving only chemical individuals) 18.02 g. or some simple multiple, such as 36.04 g. or 9.01 g., etc., is the amount of water that will react with one reacting weight of any other chemical individual; and it tells us further that water contains hydrogen and oxygen in the proportion by weight of $2 \times 1.01 = 2.02$ grams of hydrogen to $1 \times 16 = 16$ grams of oxygen. (We shall see later that it tells even more than this.)

Again, 44 grams is the reacting weight of carbon dioxide. Forty-four grams of this substance contain 12 grams of carbon (represented by C) and 32 grams of oxygen (represented by 2O). The chemical formula may be therefore written CO_2, indicating the reacting weight ($12 + 2 \times 16 = 44$ grams) and the composition by weight.

SELECTION OF REACTING WEIGHTS.

So far as the definition of reacting weights given in the statement of the law is concerned, it is obviously a matter of complete indifference what particular multiple of the experimental values given on page 120 is taken as the reacting weight of the various substances. In the case of carbon, for instance, 24 grams or 6 grams would serve as well as 12 grams, for the weights of carbon taking part in reactions with 16 grams of oxygen are simple multiples of these numbers just as much as they are of 12. Since, however, these quantities are employed as very convenient units for expressing quantitative relations, and are the basis of the system of chemical symbols, it is indispensable to have uniformity in usage among chemists, and it is necessary to fix upon particular multiples. If for any reason at all, however trivial, there is any advantage in the use of one multiple over that of another,—if, for example, some additional information about the substances can be expressed by one set of numbers better than by another set, then these factors should certainly have an influence in the choice of the multiples.

Reacting Weights of Elements.

In selecting the particular multiple of the experimental values given on p. 120 as reacting weights for the elements, chemists have been influenced by such things as specific heat, Mendelejeff's classification, isomorphism, etc. These need not be gone into in an elementary text book, as the reasons on which the selections have been based are unsatisfactory, and the additional information given by the particular values chosen is of minor importance. All that concerns us at the present day is that the choice has been made and is accepted by all chemists.

Reacting Weights or Formula Weights of Gaseous Compounds.

In the case of gaseous compounds, however, the choice is based on a very important law, and the additional information given is of great value. In this case it is possible, owing to a very striking and simple natural law connecting reacting weights and volume, to express by the chemical formula, not only the weight of a gas taking part in a reaction, but also its volume at known pressure and temperature. It is found, namely, that the *volumes* occupied by the *weights* of the gases given in the table on p. 120, bear a very simple relation to one another, at the same temperature and pressure.

28 g.	carbon monoxide	at 0° C.	and 760	mm.	occupy	22.4	litres[*]
4.01 g.	methane	"	"	"	"	5.6	"
11 g.	carbon dioxide	"	"	"	"	5.6	"
22 g.	"	"	"	"	"	11.2	"

These volumes are in the proportion of 4 : 1 : 1 : 2. By choosing suitable simple multiples of 28 grams, 4.01 grams, 11 grams etc., as the reacting weights, we can arrange that the volume filled by one reacting weight of each gas will be the same. We select $1 \times 28 = 28$ grams for carbon monoxide, $4 \times 4.01 = 16.04$ grams for methane and $2 \times 22 = 44$ grams for carbon dioxide, *thus obtaining quantities which are not only the reacting weights of the various substances, but are also the weights of the various gases which occupy a volume of 22.4 litres at 0° and 760 mm.*[†] These weights are called *formula weights*, and the symbols representing them are called the *chemical formulae*. The advantage in the choice of 22.4 litres instead of 5.6, 11.2 or 44.8, etc., is that on the one hand

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[*] The values given in this table have been calculated from experimental data given in the previous part of the book. For example, on p. 105, the weight of 1 litre of carbon monoxide was given as 1.250 g. at 0° C. and 760 mm., consequently, 28 grams fill 22.4 litres. The pupil should verify the other numbers given in the table, making use of the following data: 1 litre of carbon dioxide at 0° C. and 760 mm. weighs 1.964 g. (see p. 101), and 1 litre of methane at the same pressure and temperature 0.7161 g. (see p. 108).

It may be noted here that the generalization or "law," that the weights of 22.4 litres of all gases at 0° C. and 760 mm. are simple multiples of the reacting weights is only approximately true, and does not hold with the same degree of accuracy as the "law of reacting weights," which is one of the most exact laws of nature known.

[†] In many cases it is not possible, or not convenient, to measure the volume of a gas at 0° C. and 760 mm. pressure, but if the measurements can be carried out at any other pressure and temperature, the volume at 0° C. and 760 mm. may be calculated from the gas laws.

we obtain the simplest possible chemical formulae, and on the other hand we avoid the use of fractions. This may be illustrated by the example of carbon dioxide. Thus, if 11.2 litres were used, the chemical formula for carbon dioxide would be $C_{\frac{1}{2}}O$; if 44.8 were employed, it would be C_2O_4.

Formula Weights of Elements.

Since many of the elements exist in the gaseous form, we can use chemical formulae for these also, to represent the weight of gas occupying 22.4 litres at 0°C. and 760 mm. Since 2.02 g. hydrogen occupy 22.4 litres,* and since the symbol H stands for 1.01 g. hydrogen, the *formula weight* of hydrogen must be 2×1.01, and its formula is therefore H_2. In this way formulae have been assigned to a number of elements, for instance O_2 for oxygen, O_3 for ozone, etc.

Example of the Determination of a Chemical Formula.

The foregoing principles may be further illustrated by the following example:

Chemical formula of Methane. On p. 108 the composition of methane was given as 1 g. carbon to 0.337 g. hydrogen. We must first express these proportions in terms of the reacting weights of each element, namely, 12 grams for carbon and 1.01 grams for hydrogen.

$$1 \text{ gram of carbon } = \frac{1}{12} = 0.0833 \text{ reacting weight (R.W.) of carbon ;}$$

$$0.337 \text{ gram of hydrogen} = \frac{0.337}{1.01} = 0.333 \text{ R.W. of hydrogen.}$$

The composition of marsh gas is therefore

0.0833 R.W. of carbon to 0.333 R.W. of hydrogen.

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* 1 litre of hydrogen weighs 0.0902 g. (see p. 67).

By dividing through by the smaller number, 0.0833, we find that the proportions are

1 reacting weight of carbon to 4 reacting weights of hydrogen,

consequently, any one of the formulae CH_4, C_2H_8, $C_{\frac{1}{2}}H_2$, etc., would express the composition of marsh gas. In order to decide between these formulae we must know the weight of 22.4 litres of marsh gas at 0° C. and 760 mm. From the weight of one litre, given on p. 108 as 0.7161 gram, we can obtain the weight of 22.4 litres by simply multiplying 0.7161 by 22.4. This gives 16.04 grams as the formula weight. The chemical formula is therefore CH_4 ($12\text{ g.} + 4 \times 1.01\text{ g.} = 16.04\text{ g.}$).

Formula Weights of Non-Gaseous Compounds.

In the case of chemical compounds which are not gases, the selection of the formula weights has been influenced by the desire for simplicity of formula, and by other considerations which cannot be discussed here.

Chemical Equations.

The weights of chemical individuals entering into and produced by a reaction may be conveniently expressed by their chemical formulae arranged with the signs + and = in the form of an equation. The reaction between hydrogen and oxygen forming steam may be represented thus:

$$2\ H_2 + O_2 = 2\ H_2O.$$

Two formula weights (4.04 grams) of hydrogen react with one formula weight (32 grams) of oxygen to give two formula weights (36.04 grams) of steam. Moreover, as all three substances are gaseous, each formula weight represents the same volume of gas, namely, 22.4 litres (at 0° C. and 760 mm.), so that the equation indicates that 44.8 litres of hydrogen and 22.4 litres of oxygen react, giving a volume of steam which, calculated to 0° C. and 760 mm., would be 44.8 litres.*

REVIEW.

It will be seen from the above, that in the case of elements we have two sets of chemical units. Firstly, we have units called *reacting weights* represented by symbols, such as O for 16 g. of oxygen, H for 1.01 g. of hydrogen, etc. In many books on chemistry these reacting weights are called *atomic weights*, a name derived from Dalton's atomic hypothesis. Secondly, we have *formula weights* for those elements which can be obtained in the gaseous form. These are represented by chemical formulae, such as H_2 for 2.02 g. of hydrogen, O_2 for 32 g. of oxygen, and O_3 for 48 g. of ozone, etc. These weights are always simple multiples of the reacting weights, and are always the weight of 22.4 litres of the gas at 0° C. and 760 mm. pressure. They are often called *molecular weights*, owing to a certain hypothetical interpretation of the facts on which the law of reacting weights is based.

For chemical compounds we have only one set of chemical units; these are represented by the chemical formulae, and may be called indiscriminately reacting weights or formula weights. The formula indicates the

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* This equation, of course, does not imply that less than 16 grams of oxygen, or 2.02 grams hydrogen, cannot take part in a reaction. It must be remembered that chemical equations represent the *proportions* in which substances react, and that instead of grams we might have used tons, pounds, or even milligrams, were not the gram the customary unit of weight in science. The same may be said of the volumes of gaseous substances. This equation indicates that, expressed in any units at all, two volumes of hydrogen react with one volume of oxygen, giving two volumes of steam.

composition of the substance, its reacting weight, and, in the case of gases, the weight of 22.4 litres of the substance at 0° C. and 760 mm. Thus the formula CO_2 indicates that carbon dioxide is composed of carbon and oxygen in the proportion of 12 grams of carbon to 32 grams of oxygen; that its reacting weight is 44 grams, and that 44 grams of this gas occupies 22.4 litres at 0° C. and 760 mm.

It must not be forgotten that the law of reacting weights is applicable only to reactions in which none but chemical individuals are involved, and that many substances react in proportions which are *not* simple multiples of their reacting weights, to form solutions. For example, although oxygen and hydrogen, in forming the compounds water and hydrogen peroxide, take part only in simple multiples of 16 g. and 1.01 g. respectively, there is no such restriction on the proportions in which the two gases may be brought together to form a gaseous solution.

PROBLEMS.

1. Find, from the following data, the reacting weights (or some simple multiple of them) of the chemical substances in italics.

(*a*) 7.95 g. of *copper*, on being burned in oxygen, gave 9.95 g. of *copper oxide*.

(*b*) 1.471 g. of *potassium chlorate*, on being heated, evolved pure oxygen leaving 0.8952 g. *potassium chloride*.

(*c*) 9.34 g. *of a certain metal* combined with oxygen, giving 11.63 g. of the *oxide of this metal*.

(*d*) 16.87 g. of the *oxide of the metal M* unite with 1.982 g. of water, giving a *hydroxide*. 23.01 g. of the oxide of the same metal M unite with 2.40 g. more oxygen, giving *another oxide of M*. In this problem the reacting weight of water (18.02 g.) is assumed to be known.

2. Find the simplest formula, and the corresponding formula weight, for each of the chemical substances having the following composition, using the reacting weights given on p. **215**.

(*a*) Copper, 34.64% ; iron, 30.45% ; sulphur, 34.91%.

(*b*) Calcium, 29.45% ; sulphur, 23.55% ; oxygen, 47.01%.

(*c*) Copper, 25.47% ; sulphur, 12.83% ; oxygen, 57.66% ; hydrogen, 4.04%.

(*d*) Sodium, 21.64% ; chlorine, 33.29% ; oxygen, 45.06%.

(*e*) Nitrogen, 13.86% ; potassium, 38.61% ; oxygen, 47.52%.

3. One litre of the vapor of a volatile chemical compound, measured at 150° C. and 750 mm. pressure, is found to weigh 0.512 g. What is the formula weight of the compound?

4. The density of a gaseous chemical compound is found to be 2.80, air being taken as unity (22.4 l. of air at 0° C. and 760 mm. weigh 28.97 g.). Find the formula weight of the compound.

5. 0.69 g. of a chemical substance, when converted into a gas, is found to exert a pressure of 535 mm. when filling a volume of 320 cc. at 360° C. Find the formula weight of the gas.

6. The density of a gaseous element is found to be 13.9, compared with hydrogen as unity. The formula of hydrogen being H_2, find the formula weight of the gaseous element.

7. The formula weight of oxygen has been found to be 32 g. What will be the weight of 150 cc. of the gas at 30° C. and 745 mm.?

8. What will be the weight of 100 cc. of hydrogen chloride (formula weight 36.45 g.) at 20° C. and 755 mm.?

9. The formula of a gaseous compound is found to be C_3H_8. What will be the weight of a litre of the gas at 50° C. and 770 mm. pressure?

10. At what temperature will 15 grams of ammonia (NH_3) occupy 25 litres, under a pressure of 730 mm.?

11. Calculate the formulae of the following substances. As they are all gases, a knowledge of their density is necessary in order to fix the formulae. The compositions of the gases are given in parts by weight.

(*a*) The composition of the substance is 92.24% carbon and 7.76% hydrogen. One litre of the gas weighs 1.161 g. at 0° C. and 760 mm.

(*b*) The composition of the substance is 27.27% carbon and 72.73% oxygen. Ten litres of the gas at 170° C. and 750 mm. weigh 18.25 g.

(*c*) The composition of the substance is 92.24% carbon and 7.76% hydrogen. At 127° C. and 735 mm. pressure 0.5 litre weighs 1.150 g.

(*d*) The composition of the gas is found to be 82.25% nitrogen and 17.75% hydrogen. Its density is 0.589, taking air as unity. 22.24 litres of air at 0° C. and 760 mm. weigh 28.97 g.

(*e*) The composition of the substance is 87.42% nitrogen and 12.58% hydrogen. One gram of the gas at 77° C. and 745 mm. pressure occupies a volume of 0.9118 litre.

12. At 17° C. and 750 mm. pressure, 26.51 cubic centimetres of a gas containing carbon and hydrogen only, on being exploded with oxygen, gave 0.0484 g. carbon dioxide and 0.0396 g. water. Find the formula of the gas. Calculate also the equation expressing the reaction, and the volume of oxygen used up.

13. 1.5429 g. of mercuric oxide (HgO), when heated, gave 1.4320 g. of mercury (Hg) and 86.2 cc. of oxygen measured at 20° C. and 760 mm. Calculate the chemical equation representing the reaction.

14. 1.3622 g. of potassium chlorate ($KClO_3$), when heated, gave 0.8289 g. of potassium chloride (KCl) and 373.3 cc. of oxygen measured at 0° C. and 760 mm. Calculate the chemical equation expressing the reaction.

15. 11.449 g. of a certain gas burns, using up 35.20 g. of oxygen, and forming 38.72 g. carbon dioxide and 7.929 g. water. One litre of the gas at 0° C. and 760 mm. weighs 1.161 g. Calculate the equation representing the reaction.

16. The change of limestone into lime and carbon dioxide is found to take place according to the following equation :—

$$CaCO_3 = CaO + CO_2.$$

If 1000 g. limestone were used, calculate the weight of lime obtained, and the volume of carbon dioxide produced, when measured at 17° C. and 745 mm.

17. The action of zinc on hydrochloric acid is represented by the equation

$$Zn + 2\,HCl = H_2 + ZnCl_2.$$

How much zinc would be necessary to produce 30 litres of hydrogen, measured at 27° C. and 760 mm.?

18. Marble reacts with hydrochloric acid, giving carbon dioxide as follows :—

$$CaCO_3 + 2\,HCl = CaCl_2 + H_2O + CO_2.$$

What weight and what volume of carbon dioxide would be formed from 50 grams of marble at 17° C. and three atmospheres pressure?

19. The formula weight of a gas is found to be represented by the formula C_2H_6. Calculate its density, taking (*a*) air as unity (22.4 l. of air at 0° C. and 760 mm. weigh 28.97 g.), (*b*) hydrogen as unity, the formula of hydrogen being H_2.

CHAPTER XII.—CHLORINE AND ITS COMPOUNDS

CHLORINE.

Supplies: Apparatus as shown in Fig. 30, manganese dioxide, hydrochloric acid, bottles for collecting gas, powdered antimony, magnesium ribbon, hydrogen generating apparatus, turpentine, a candle, phosphorus, sodium, deflagrating spoon, filter paper, pieces of colored calico, solutions of logwood, litmus, indigo, potassium permanganate.

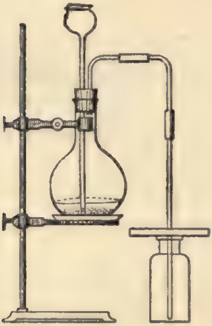

FIG. 30.

Preparation.* Arrange apparatus as shown in Fig. 30. In the flask place about 30 grams of manganese dioxide and pour in enough hydrochloric acid to cover it well. Shake the flask to prevent the dioxide adhering to the bottom, heat the sandbath gently, and collect several bottles of the gas by downward displacement of air. (See also chlorine water, p 138.)

What is the color of the gas? Has it an odor? Is it lighter or heavier than air? Does it burn or support combustion? What effect has it on moistened litmus?

Experiments with Chlorine:

1. Generate some hydrogen,† and after testing care-

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* Chlorine should be prepared in a hood or fume cupboard, and care should be taken not to inhale the gas, as it has a very irritating effect on the sensitive tissue of the throat and lungs.

† Illuminating gas may be used for this experiment instead of hydrogen, but in this case a black smoke will be formed owing to the presence of carbon in the illuminating gas.

fully to make sure that all air has been expelled, light a jet and introduce it into a jar containing chlorine, as shown in Fig. 31.

Does the hydrogen continue to burn? Is heat given out by the reaction? What change is seen to take place in the gas in the bottle? Has the new gas the color or odor of chlorine?

Add some water, shake, and test with litmus. To what class of substances does the solution belong?

2. Lower a lighted candle slowly into a bottle filled with chlorine. What do you observe?

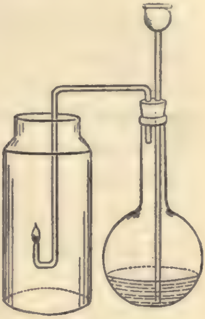

FIG. 31.

Test the gas in the bottle with litmus. Does chlorine react readily with carbon?

3. Saturate a piece of filter paper with fresh turpentine and drop it into a bottle filled with chlorine.

This is an example of spontaneous combustion in which oxygen takes no part (see p. 41). The wax or tallow of the candle, and the turpentine, contain carbon and hydrogen. The hydrogen reacted with the chlorine, leaving carbon.

4. Into another bottle containing chlorine sprinkle a little powdered antimony.

5. Ignite a piece of magnesium ribbon and hold it in a bottle filled with chlorine gas.

6. Into a deflagrating spoon put a piece of freshly cut phosphorus and lower it into a bottle filled with chlorine.

7. Into another bottle of chlorine place a colored flower, or a piece of moistened colored calico.

8. Pour a small quantity of a solution of logwood, litmus, or indigo into a bottle containing chlorine.

The gas formed in Exp. 1, p. 134, was a *chemical compound* containing only hydrogen and chlorine, and called hydrogen chloride. If this experiment had been carried out quantitatively, or if the substance had been decomposed and the weights of the products determined, hydrogen chloride would have been found to "contain" 2.77% of hydrogen and 97.23% of chlorine. From this we can get the reacting weight of chlorine and of hydrochloric acid (or some multiple of them) as follows: --

Since 2.77 g. hydrogen + 97.23 g. chlorine, give 100 g. hydrogen chloride, 1.01 g. hydrogen + 35.45 g. chlorine, give 36.46 g. hydrogen chloride. Since these substances, by the law of reacting weights, must react in proportion to simple multiples of their reacting weights, 35.45 g., or some simple multiple of it, must be the reacting weight of chlorine, and 36.46 g., or some simple multiple, the reacting weight of hydrogen chloride. 35.45 g. itself has been chosen as the reacting weight of chlorine, for reasons referred to on p. 124, and is represented by the symbol Cl.

From the density of chlorine gas, its formula weight and its formula can be now determined. (Read pp. 124–127 again carefully). 22.4 litres of chlorine at 0° C. and 760 mm. are found to weigh 70.9 grams. This is therefore taken as its formula weight; and since Cl stands for 35.45 g., the formula must be written Cl_2.

In the same way, the formula of hydrogen chloride may be found. 22.4 litres of the gas are found to weigh 36.46 g. We therefore write the formula HCl (since 1 01 + 35.45 = 36.46) and not H_2Cl_2, or H_3Cl_3.

The pupil is again reminded here of the meaning of such formulae. HCl expresses the following facts: (1) that the composition by weight of hydrogen chloride is 1.01 g. of hydrogen to 35.45 g. of chlorine; (2) that 36.46 is the reacting weight of hydrogen chloride, *i.e.*, that in any chemical reaction between this substance and another chemical substance, some simple multiple of 36.46 g. will react with the reacting weight of the other substance. For instance, 2 × 36.46 or 72.92 g. of it are found to react with 56.1 g. of calcium oxide, which, as explained on p. 120, is the reacting weight of this substance; (3) that 22.4 l. of hydrogen chloride at 0° C. and 760 mm. weigh 36.46 grams.

The reaction of Exp. 1 may now be expressed in the form of an equation, thus:

$$H_2 + Cl_2 = 2\ HCl.$$

9. Repeat experiment 6, using, instead of phosphorus, a piece of sodium the size of a pea. Heat it in the flame of a spirit lamp until it ignites and lower it into the bottle filled with chlorine. Add a few cubic centimetres of water and shake. Pour this into a dish and evaporate to dryness.

What is the appearance of the residue? Taste it.

The water was added simply to dissolve the solid from the sides of the jar where it has settled. The solution on evaporation left the dry salt again. The substance formed is common salt. It contains only sodium and chlorine, and had the experiment been carried out quantitatively, it would have been found that 1 g. of chlorine had reacted with 0.65 g. sodium, forming 1.65 g. of common salt. Common salt is a substance of constant composition or a chemical compound, consequently, the law of reacting weights will apply to this reaction. The reaction is expressed by the equation—

$$2\ Na + Cl_2 = 2\ NaCl,$$

in which Na represents 23.05 g. of sodium, and NaCl $23.05 + 35.45 = 58.50$ g. of common salt. Since sodium chloride is not a gas, its formula conveys no information as to the weight of 22.4 litres. NaCl is used rather than Na_2Cl_2, etc., for the sake of simplicity.

A calculation should be made to show that this equation represents the experimental proportions given above.

Experiments 4, 5 and 6 also involve only chemical individuals, and therefore may be expressed by chemical equations. Determinations of the weights and volumes of the materials involved have been made and the equations found to be as follows :—

Antimony and chlorine give antimony pentachloride.

$$2\ Sb + 5\ Cl_2 = 2\ SbCl_5.$$

Magnesium and chlorine give magnesium chloride.

$$Mg + Cl_2 = MgCl_2.$$

Phosphorus and chlorine give phosphorus trichloride.

$$2\ P + 5\ Cl_2 = 2\ PCl_5.$$

What quantitative information is expressed by each equation? Refer to the table of symbols and reacting weights, p. 215.

Chlorine Water.

Prepare chlorine as before, but allow the gas to bubble for some time through cold water contained in a large bottle. (Why cold water? See p. 100.)

Note the gradual change in color of the liquid. Is its composition constant? Is chlorine water a solution or a chemical compound?

Experiments with Chlorine Water :

10. Fill a flask with chlorine water, support it as shown in Fig. 32, place it in the sunlight and allow it to remain for a day or two.* Invert the flask, remove the cork and insert a glowing splinter. Test the liquid in the flask with litmus.

What are the products of the reaction?

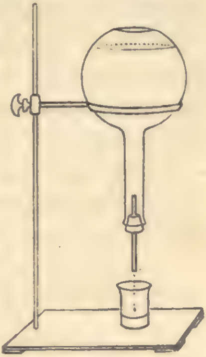

FIG. 32.

11. Pour some chlorine water into solutions of logwood, litmus, indigo, and potassium permanganate.

12. Drop a piece of colored calico into chlorine water.

13. Write on a piece of printed paper with ordinary ink, place it in chlorine water and allow it to remain for some time.

Printer's ink is composed largely of lampblack, while writing ink generally contains an organic coloring matter. Refer to Exp. 2, and account for the different action of the chlorine on the two kinds of ink.

Notes on Chlorine.

Chlorine is a greenish yellow gas, having a disagreeable, suffocating odor. Its reacting weight is 35.45 g., and its formula Cl_2, indicating that it is about 2.5 times heavier than air, 22.4 l. of which weigh 28.97 g.

Chlorine does not burn in air, but many substances burn in it, forming compounds of chlorine. When these compounds contain only chlorine and *one* other

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* For other examples of the action of light, see pp. 17, 22 and 23.

element, they are called *chlorides*. This nomenclature is analogous to that employed for compounds of oxygen with *one* other element, to which, it will be remembered, the term *oxides* was applied. The following are examples of chlorides :—

Sodium chloride, $NaCl$, potassium chloride, KCl, the iron chlorides, $FeCl_2$ and $FeCl_3$, copper chloride, $CuCl_2$, the phosphorus chlorides, PCl_3 and PCl_5, silver chloride, $AgCl$, etc.

Many substances containing hydrogen and carbon, such as hydrocarbons, wax, tallow, turpentine, benzene, etc., burn in chlorine with a sooty flame, the hydrogen combining with the chlorine, forming hydrochloric acid, while the carbon is liberated as a black smoke.

Chlorine is soluble in about one-third its volume of cold water, giving a solution which has a corrosive and bleaching action on many substances. When a vessel containing the solution is left open, chlorine is given off. (See soda water, p. 100.) This is why chlorine water always has the odor of chlorine.

When the solution is exposed to the sunlight a gradual change takes place, the chlorine forming hydrochloric acid with the hydrogen of the water, while the oxygen is given off. A large number of other reactions are promoted by light, many of which form the basis of photographic processes. The changes which take place in the leaves of plants during growth also belong to this class of reactions.

When water and chlorine are in contact with substances which can unite with oxygen, a reaction takes place even without the aid of light, and it is to this reaction that the bleaching properties of chlorine are due. When certain coloring matters are brought into contact with chlorine water, or when a colored material is moistened with water and placed in chlorine gas, hydrochloric acid is formed from the chlorine and hydrogen of the water, while the oxygen combines with the coloring matter, forming uncolored products. The value of chlorine as a *disinfectant* and *deodorizer* is due to the same cause.

It must not be supposed that oxygen is first set free from the water and chlorine, and then combines with the coloring matter. If this were the case, bubbles of the gas would probably be seen; many dye stuffs, moreover, may be bleached by chlorine water, which are not acted on by oxygen.

HYDROGEN CHLORIDE.

SUPPLIES: Apparatus as shown in Fig. 33, common salt, sulphuric acid, bottles for collecting gas, splinters, ammonium hydroxide, litmus paper, filter paper, candle.

Preparation. In the test tube, Fig. 33, place about 20 g. of common salt, and pour down the safety tube enough sulphuric acid (slightly diluted) to cover it. Heat gently, if necessary, and collect several bottles of the gas.

Describe the gas obtained.

Suggest a possible method for preparing chlorine from common salt. (See Fig. 30, p. 134.)

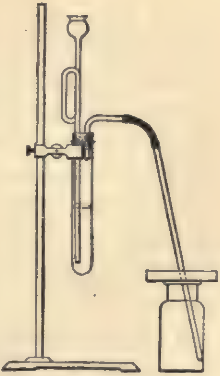

FIG. 33.

Experiments with Hydrogen Chloride:

1. Hold a burning splinter in a bottle filled with the gas. Lower into it a burning candle.

2. Hold a moistened piece of litmus paper in the mouth of the same bottle of gas.

3. Moisten a piece of filter paper with ammonium hydroxide and insert it into the mouth of the same bottle.

4. Devise an experiment to ascertain whether the gas is heavier or lighter than air.

5. Place another bottle filled with the gas, mouth downwards in a trough of water, and shake gently. Slip a piece of paper under the mouth of the bottle, invert it, and test the contents with litmus paper. Taste the liquid.

Compare the solubility of this gas with that of carbon dioxide. (See Exp. 3, p. 96.)

Notes on Hydrogen Chloride.

Hydrogen chloride is a colorless gas having a choking, pungent odor. When it escapes into moist air it forms a liquid solution with the water vapor in the atmosphere, the drops of which are visible to the eye as a white cloud. It is heavier than air and does not burn or support combustion. It is very soluble in water, 1 litre of water at ordinary temperature dissolving 450 l. of the gas. The solution thus formed is an acid* called hydrochloric acid.

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* It has been the custom with some chemists to extend the term *acid* so as to include some materials which, though themselves possessing no acid properties, give acids on solution in water. Thus, in some books on chemistry, the gases hydrogen chloride and carbon dioxide are called hydrochloric acid and carbonic acid respectively. As, however, this nomenclature has not been carried out consistently, the term will be restricted in this book as far as possible to solutions which turn blue litmus red.

HYDROCHLORIC ACID.

SUPPLIES: Hydrochloric acid, baking soda, calcium carbonate (limestone or marble), ammonium carbonate, sodium hydroxide, magnesium, zinc, quicklime, sulphuric acid.

Experiments with Hydrochloric Acid:

1. Put a little dilute hydrochloric acid in a test tube, and to it add some baking soda.

2. Repeat the experiment with some limestone and magnesium carbonate

Repeat also with a strip of magnesium and a piece of zinc; test the gas given off with a lighted match.

3. To about 10 cubic centimetres of dilute hydrochloric acid in a dish, add washing soda slowly, in small portions, till effervescence ceases. Taste the solution obtained. Evaporate it to dryness and compare the product with that from sodium and chlorine, p. 137.

4. To the same amount of acid as was used in the last experiment add slowly a solution of sodium hydroxide till a drop of the liquid just begins to turn red litmus blue. Note roughly the amount of sodium hydroxide solution necessary for this purpose. Evaporate the solution until a moist mass of crystals is obtained. Remove the crystals and dry *thoroughly*, by pressing between filter paper. Taste the dry crystals and compare them with those obtained in experiment 3.

Some of the class should repeat the experiment, using twice as much sodium hydroxide as before, and the rest of the class should use the same amount of hydroxide with about twice as much acid. In both cases the crystals should be dried carefully and compared with those previously obtained.

5. Add quicklime to a little hydrochloric acid and see whether the acid properties can be made to disappear.

6. Repeat experiments 1 and 2, using sulphuric acid instead of hydrochloric.

Properties of Acids.

The above experiments show that the acids investigated, besides turning blue litmus red, have also other properties in common. They both effervesce with carbonates, giving a gas that might be shown to be carbon dioxide (How?); both effervesce with metals giving off hydrogen, and both have a sour taste.

Formation of Salts from Acids.

Another common property illustrated by the experiments is the ability of acids to form salts (see also p. 98). Experiments 3 and 4 showed that the *same* product (common salt) was formed by two different methods. Experiment 4 showed that the same substance was produced in spite of the difference in the proportions of acid and base used. It is consequently a chemical compound, and not a solid solution nor a mixture (see page 116). (Compare the formation of salt by this method with the solidification of brass from molten zinc and copper. In the latter case the composition, and all other properties, of that part of the mass which has solidified to brass, vary with the proportion of zinc and copper in the molten part.)

There are many ways of preparing salts; the most general are by the action of:

1. Acids on bases, *e.g.*, sodium hydroxide.
2. Acids on metals, *e.g.*, magnesium.
3. Acids on metallic oxides, *e.g.*, calcium oxide.
4. Acids on carbonates, *e.g.*, sodium carbonate.

Use of Chemical Equations for Reactions between Solutions.

The neutralization of hydrochloric acid by a solution of sodium hydroxide in experiment 4, p. 143, gave a new solution with neither acid nor basic properties. Since the substances involved are not chemical individuals,* this reaction, strictly speaking, cannot be expressed by a chemical equation at all. It is, nevertheless, customary to say that it takes place according to the equation :

$$HCl + NaOH = NaCl + H_2O.$$

This use of chemical equations is quite different from that explained on page 127.

The formula HCl stands for 36.46 grams of hydrogen chloride ; NaOH for 40.06 grams of sodium hydroxide or caustic soda ; NaCl for 58.50 grams of common salt ; and H_2O for 18.02 grams of water, ice, or steam, and the equation just written represents quantitatively the changes which take place when hydrogen chloride is brought together with sodium hydroxide, forming solid salt and steam. The *solutions* used in experiment 4 are not represented at all.

The same quantity of salt and water, however, might be obtained from hydrogen chloride and sodium hydroxide as the net result of the following series of operations, involving four distinct reactions, which, though they may look complicated on paper, are easily and quickly carried out in the laboratory.

(i) 36.46 grams of hydrogen chloride are dissolved in, say, 1000 grams of water, forming 1000 + 36.46 grams of hydrochloric acid.

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* They are *solutions* of chemical individuals.

(ii) 40.06 grams of sodium hydroxide are dissolved in, say, 1000 grams of water, forming 1000 + 40.06 grams of sodium hydroxide solution.

(iii) The two solutions are brought together, when they form 2000 + 36.46 + 40.06 grams of brine.

(iv) The brine is evaporated to dryness, giving 58.50 grams of common salt and 2000 + 18.02 grams of steam.

The water used in operations i and ii is thus recovered in operation iv, with 18.02 grams additional; and the net result of all four reactions is the same as that of the single reaction that takes place when the 36.46 grams of hydrogen chloride act on the 40.06 grams of sodium hydroxide. Thus, although

$$HCl + NaOH = NaCl + H_2O$$

is really the chemical equation for the single reaction, and does not represent any one of the four reactions in which solutions are involved, yet the quantities represented in it are those actually lost and gained as a result of four reactions, and in this sense the equation may be said to represent them. Quantitative experiments have shown that the following equations may be employed in this way, the water used in making the solutions being ignored. Thus:—

Hydrochloric acid + magnesium oxide = magnesium chloride + water.
$2\,HCl + MgO = MgCl_2 + H_2O.$
Hydrochloric acid + calcium oxide = calcium chloride + water.
$2\,HCl + CaO = CaCl_2 + H_2O.$
Hydrochloric acid + sodium carbonate = sodium chloride, carbon dioxide and water.
$2\,HCl + Na_2CO_3 = 2\,NaCl + CO_2 + H_2O.$

Notes on Hydrochloric Acid.

Commercial hydrochloric acid is a concentrated solution of hydrogen chloride in water. The solution, saturated at ordinary pressure and temperature, contains about 42% by weight of the gas, while ordinary commercial acid contains about 25%. The pure acid is colorless, but the cheaper commercial preparations have a yellow color, due to impurities. Large quantities of the acid are used for preparing chlorine, for the manufacture of bleaching powder, and for making the chlorides of the metals.

Test for Hydrochloric Acid or any dissolved Chloride. To a solution of any chloride add a few drops of a solution of silver nitrate. Divide the precipitate; to one part add nitric acid, and to the other ammonia solution.

Try the same experiment with potassium *chlorate*, and with the residue after heating some potassium chlorate.

Do solutions of *all* substances containing chlorine answer this test? How would you distinguish a solution of sodium chloride from one of hydrochloric acid, and from one of potassium chlorate?

Bleaching Powder.

Manufacture. Bleaching powder is manufactured by the action of chlorine gas on slaked lime. Quicklime is carefully slaked with water, forming slaked lime or calcium hydroxide. This powder is sifted into a large chamber lined with lead or tarred brick, until the floor is covered to a depth of three or four inches. The chlorine enters at the top and settles slowly to the floor, where it is absorbed by the lime.

Bleaching Action of Bleaching Powder. To a small quantity of bleaching powder (chloride of lime) add sufficient water to make a thin paste. Immerse a strip of colored calico in this paste until it is thoroughly saturated. Then place it in a dilute solution of sulphuric acid, hydrochloric acid, or vinegar. Remove the calico and wash with water. Repeat the process as often as may be necessary.

Bleaching powder, usually termed chloride of lime, is used in large quantities to whiten cotton and linen goods and paper pulp. It is also used as a disinfectant. In both cases its action is due to the chlorine which is easily liberated from it. When dry it is a white powder, but when exposed to air it gives off chlorine, and changes to a deliquescent substance containing calcium chloride.

Other Compounds of Chlorine.

Besides the compounds of chlorine with hydrogen, sodium, etc., already described, there are many in which *two* or *more* other elements are combined with the chlorine. One example is potassium chlorate, a substance already mentioned in connection with the preparation of oxygen. As will be remembered, this substance, on being heated, turned into oxygen and potassium chloride. The latter salt has the composition shown by the formula KCl, or 39.15 g. potassium to 35.45 g. chlorine, corresponding to the reacting weight 74.60, as would have been found in experiment 11, p. 47, had the work been done accurately.

1 g. potassium chlorate gives 0.3915 g. oxygen
+ 0.6085 g. potassium chloride;
or, 122.6 g. potassium chlorate gives 48 g. oxygen
+ 74.60 g. potassium chloride.

Now, 74.60 g. of potassium chloride is represented by KCl, 48 g. (= 3 + 16) of oxygen by 3 O, consequently, the 122.6 g. of potassium chlorate must be represented by $KClO_3$. This is taken as the formula of the substance, and the equation is written

$$2\ KClO_3 = 2\ KCl + 3\ O_2.$$

the amounts being doubled in order that the *formula* of oxygen gas, O_2, may be used without introducing fractions before the symbols.

FORMULAE OF ACIDS.

As already seen, we have a number of solutions called acids, with many common properties. These are not chemical individuals, and cannot be expressed by chemical formulae. (See footnote, p. 142). In many cases, however, chemical compounds which contain hydrogen may be obtained by evaporating such solutions, and in these cases the solution is said to "contain" that chemical substance. In other cases it is not possible to obtain from acid solutions chemical substances containing hydrogen. Since, however, it is desirable to indicate by the formulae the similarity in properties of acid solutions, it has been found convenient to speak of all acid solutions as if they "contained" *compounds* of hydrogen to which chemical formulae are assigned. The following illustrations will make this clear:—

A solution of carbon dioxide, CO_2, in water, has acid properties, but no chemical compound containing hydrogen can be obtained from this solution. We say, however, that the solution contains a substance with the formula H_2CO_3, thus suggesting by the H in the formula the similarity of carbonic acid to hydrochloric acid. That this will not lead in practice to any erroneous results will be seen, when it is noted that the formula H_2CO_3 might be written $CO_2 + H_2O$, so that the formula represents the composition of a substance which contains one reacting weight of carbon dioxide to one reacting weight of water, and neglects all the rest of the water, just as in expressing the reaction between sodium hydroxide and hydrochloric acid, the water used as a solvent was ignored. (See p. 146.)

The reaction between sodium hydroxide solution and carbon dioxide solution (carbonic acid) may be accordingly expressed as follows :—

$$2\,NaOH + H_2CO_3 = Na_2CO_3 + 2\,H_2O\,;$$

$$\text{or, } 2\,NaOH + CO_2 = Na_2CO_3 + H_2O.$$

In the same way the acid, which, with potassium hydroxide, gives potassium chlorate solution, is said to contain a substance represented by the formula $HClO_3$, though no chemical substance of this composition has ever been obtained. This solution is called chloric acid, and is the acid corresponding to a number of salts termed chlorates, of which potassium chlorate ($KClO_3$) is the best known. How might pure potassium chlorate be made from chloric acid ?

Compounds of Chlorine Continued—Nomenclature.

There are several other sets of salts containing chlorine, namely, chlorites, corresponding to chlorous acid ; hypochlorites, corresponding to hypochlorous acid ; and perchlorates, corresponding to perchloric acid. The formulae of the acids, and of their potassium salts, are given in the following table :—

Perchloric acid,	$HClO_4$	Potassium perchlorate,	$KClO_4$
Chloric acid,	$HClO_3$	Potassium chlorate,	$KClO_3$
Chlorous acid,	$HClO_2$	Potassium chlorite,	$KClO_2$
Hypochlorous acid,	$HClO$	Potassium hypochlorite,	$KClO$

With the exception of potassium chlorate, these substances are of no great importance, but are mentioned here merely to illustrate the system of nomenclature adopted in chemistry.

The best known acid in a series is generally given a name terminating in *-ic*. If there is another acid containing a *larger* proportion of oxygen, the syllable *per-* is prefixed to the name of the common acid. For an acid with a *smaller* proportion of oxygen than is contained in the first acid, a name terminating in *-ous* is employed, while the prefix *hypo-* is employed to indicate an acid containing a *still smaller* proportion of oxygen.

In designating the salts corresponding to these acids, the termination *-ate* is used when the name of the acid ends in *-ic*, and the termination *-ite* when the acid ends in *-ous*.

VALENCY.

The composition of a substance can be found only by experiment. The law of reacting weights, however, helps us to remember this composition, because we know that only simple multiples of certain weights, namely, the reacting weights of the elements, are involved. These reacting weights, all of which were, of course, determined experimentally, are recorded in a short table for reference, p. 215.

Taking the composition of magnesium chloride as an example, the tables give the reacting weight of magnesium to be 24.36, and that of chlorine to be 35.45, so that it is only necessary to remember that the multiples involved are *one* for magnesium and *two* for chlorine, as expressed in the formula $MgCl_2$. It would be difficult to remember these multiples in all the innumerable chemical compounds, were it not for a simple natural law, known as the valency principle. We shall first explain what is meant by the valency of an element and then give examples to illustrate the law.

Hydrogen and chlorine are said to be monovalent elements, or to have the valency *one*. On examining the formulae of a number of compounds of these elements, such as :—

KCl
$NaCl$
$CaCl_2$
$MgCl_2$
PCl_3
$SbCl_3$
CCl_4
H_2O
NH_3
CH_4

we see that one reacting weight of the other element combines with one, two, three, or four reacting weights of chlorine, or hydrogen as the case may be. We say that those elements such as potassium and sodium, one reacting weight of which combines with one reacting weight of chlorine, or of hydrogen, are monovalent, or that their valency is *one*. Those like calcium, magnesium and oxygen, one reacting weight of which combines with *two* reacting weights of chlorine, or of hydrogen, are said to be divalent. For a similar reason phosphorus, antimony and nitrogen are said to be trivalent, and carbon tetravalent.

Now it has been found, in a great many cases, that when two elements of the same valency combine, they do so in the proportion of one reacting weight of one to one reacting weight of the other, as HCl, MgO, CaO; that *two* reacting weights of a monovalent element combine with one reacting weight of a divalent element, such as Na_2O, K_2O; that *three* reacting weights of a divalent element combine with two reacting weights of a trivalent element as P_2O_3, Sb_2O_3, and so on, or in general, that if the composition of a compound be represented by the formula A_xB_y, x is to y as the valency of B is to the valency of A.

The fact that many elements have more than one valency, detracts from the value of the above principle as an aid to the memory. Phosphorus, for instance, forms two chlorides, PCl_3 and PCl_5; that is to say, the valency of phosphorus is three in the first case and five in the second, and as this double valency is characteristic for many other compounds of phosphorus, *e.g.*, P_2O_3, and P_2O_5, it is necessary to remember both valencies. In the case of some other elements the application of the principle is even less satisfactory.

The following table shows the valencies of the commoner elements, and will be found a useful aid in remembering formulae :—

TABLE OF VALENCIES

GENERALLY CONSTANT VALENCY			
Monovalent	Divalent	Trivalent	Tetravalent
Hydrogen	Calcium	Boron	Carbon
Sodium	Strontium	Aluminium	Silicon
Potassium	Barium		
Silver	Magnesium		
Chlorine	Zinc		
Bromine	Copper		
Iodine	Nickel		
	Lead		
	Oxygen		

VARIABLE VALENCY			
Element	Valency	Element	Valency
Nitrogen	3, 5	Cobalt	2, 3
Phosphorus	3, 5	Mercury	1, 2
Arsenic	3, 5	Tin	2, 4
Antimony	3, 5	Chromium	3, 6, etc.
Bismuth	3, 5, 2	Sulphur	2, 4, 6
Iron	2, 3	Manganese	2, 7, etc.

PROBLEMS.

1. What will be the probable formulae of silver oxide, silver chloride, calcium chloride, magnesium chloride, carbon chloride, aluminium chloride, lead oxide, aluminium oxide, lead chloride, copper chloride, copper oxide, zinc oxide, zinc chloride, two chlorides of iron, two oxides of phosphorus, two chlorides of mercury?

2. How many grams of hydrogen chloride would be necessary to make 50 grams of 12% hydrochloric acid?* What volume of the gas measured at 20° C. and 740 mm. would be necessary?

3. How many grams of a 10% hydrochloric acid are necessary to react with a solution of sodium hydroxide to form a solution containing 100 grams of common salt? How many grams of solid sodium hydroxide would be necessary to just neutralize the acid used?

4. How many grams of hydrogen would be produced by 25 grams of a 20% hydrochloric acid, acting on zinc, the reaction being represented by the equation:—

$$Zn + 2\,HCl = ZnCl_2 + H_2?$$

What volume of hydrogen at 17°C. and 745 mm. would be produced? How much zinc would be used up? What weight of zinc chloride would be obtained by evaporating the solution?

5. The specific gravity of a 20.3% hydrochloric acid is 1.1. Find how many cubic centimetres of this solution are necessary to neutralize 15 grams of sodium hydroxide, NaOH. How many grams of common salt could be obtained from the solution?

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* By a 12% hydrochloric acid is meant a solution containing 12 grams of hydrogen chloride to every 100 grams of solution.

CHAPTER XIII.—AIR

In Chapter II. we have seen that air contains oxygen and another gas which will not support combustion. Exp. 9, p. 33, showed that this latter gas constituted about four-fifths of the total volume of air. It consists chiefly of nitrogen—an element which will be taken up in the next chapter.

The following experiments show other methods of determining roughly the amount of oxygen in air.

Supplies: Bell jar fitted with cork, evaporating dish, phosphorus, iron filings, pyrogallic acid, caustic potash, ferrum redactum, mica, small wire tripod, calcium chloride, U-shaped tube, fresh lime water, a graduated tube fitted at one end with a stopcock (burette), a retort stand, clamp and tall glass jar.

Experiments:

1. **Phosphorus and Air.** Place a piece of phosphorus in an evaporating dish, float it on water in a trough, ignite the phosphorus and place over it the bell jar, Fig. 34. Allow the jar to remain until the gas inside becomes clear.

Account for the various changes that occur.

Depress the jar, until the water stands at the same height inside as out, remove the cork and thrust into the jar a lighted splinter.

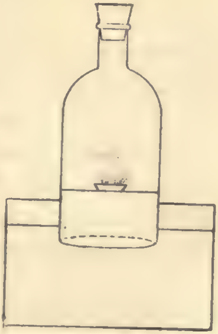

Fig. 34.

2. **Iron Filings and Air.** Wet the inside of a bottle,

sprinkle it with clean iron filings, place it mouth downwards in a dish of water and allow it to remain for a day or two.

Does the substance formed resemble rust (oxide of iron)? Which constituent of the air has been removed? Roughly, what volume?

3. **Ferrum Redactum and Air.** The last experiment may be modified as follows: Put some ferrum redactum on a small sheet of mica supported on a small wire tripod. Place both on a cork floating on water in a dish, ignite the iron and place over it a large beaker, Fig. 35. The residual gas from the last experiments, as stated above, consists mainly of nitrogen. Air contains, however, other substances besides oxygen and nitrogen, as will be shown by the following experiments :—

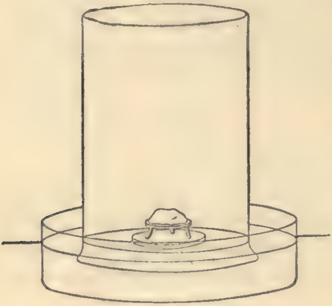

FIG. 35.

4. **Calcium Chloride and Air.** Nearly fill a U-shaped tube with small pieces of calcium chloride, and force air through it with a bellows, or a rubber bulb.

What change takes place in the chloride?

Leave some particles of calcium chloride exposed to the air in a dish for a few hours. (See deliquescence, p. 81.)

5. **Lime Water and Air.** About half fill a beaker with freshly made lime water and blow air through it with a bellows.

The simpler experiment described under carbon dioxide (Exp. 8, p. 97) may be substituted for this, if a bellows is not available.

What do experiments 4 and 5 show to be present in air?

6. **Percentage Composition of Air.** This experiment, though not differing in principle from Exps. 1, 2 and 3, allows greater accuracy in the determination of the volumes.

Support a graduated glass tube (a burette), the total volume of which is known, fitted at one end with a stopcock, as shown in Fig. 36. Fasten a piece of phosphorus, *under water*, to the end of a piece of wire, thrust it almost to the top of the tube, open the stopcock so that the water may stand at the same height inside the tube as out, then close it and read the volume of the air. Allow it to stand for a day.

Depress the tube to bring the water inside the tube to the level of the water outside, and again take the reading.

Why should the water be at the same level inside the tube as outside when the volume is measured?

What changes in the atmosphere would cause a change in the volume of the gas in the tube?

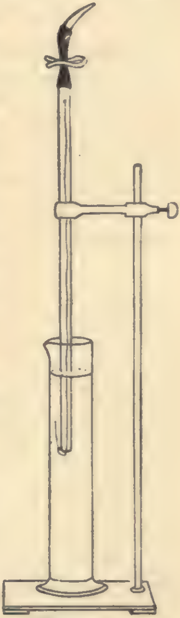

FIG. 36.

From the results of the experiment calculate the percentage by volume* of oxygen and nitrogen in the air, neglecting the carbon dioxide and water vapor, which form only a small proportion of the total volume.

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* When oxygen and nitrogen are brought together, the volume of the gas formed is the sum of the volumes of the oxygen and nitrogen, all being measured and at the same temperature and pressure.

It is evident from the above experiments that air contains several constituents. It must be, therefore, either a mechanical mixture, a chemical compound, or a solution. (See Chapter X). If air is shaken up with (say) 100 cc. of water until the latter is saturated, 0.70 cc. of oxygen will be found to have dissolved. If, on the other hand, the solubility of oxygen be determined under the same conditions, 3.5 cc. will be found to have dissolved. This shows that oxygen in the air has not the same properties as pure oxygen, and therefore that air is not a mechanical mixture of oxygen and nitrogen, according to the definition given on p. 114. (Compare the behavior of nitre in gunpowder, p. 115.) It is consequently a solution, or a chemical compound. The fact that the proportions of oxygen and nitrogen in air do not strictly accord with the law of reacting weights makes it unlikely that it is a chemical compound.*

Although it is clear that air is not a mechanical mixture in the scientific sense, nevertheless the expressions "mixing gases," and "a mixture of gases," are in such general use that it is sometimes convenient to employ them, provided no confusion can arise from this use of the terms.

Moreover, it must be pointed out that there is this difference between gaseous solutions, and liquid or solid solutions. The boiling point of a solution of alcohol and water, for example, cannot be predicted by means of any simple law, from the boiling points of the constituents, and the same is true of many other properties of the solution. In the case of gaseous solu-

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* In the case of gases, the distinction between chemical compounds and solutions cannot be based on constancy or variability of composition. The question as to why certain gases are classed as chemical compounds need not be discussed in an elementary text-book. (See footnote, p. 117.)

tions, however, all the properties may be predicted by a very simple rule. In the example given above, it will be noticed that the amount of oxygen dissolved from the air is just one-fifth of the solubility of pure oxygen, corresponding to the fraction of oxygen in air, by volume. This very simple relation between composition and properties is found to hold, approximately, for all gaseous solutions.

Another important point to be noted in connection with gaseous solutions is the following: When two gases are brought together, *in any proportion whatever*, a homogeneous solution may be formed. This is not, in general, the case when liquids or solids are involved. Although in some cases, liquids—alcohol and water, for example—may dissolve one another in unlimited quantities, these form a comparatively small class. The phenomenon usually observed is analogous to the solution of copper sulphate in water. When crystals of this substance are shaken up with water, a solution is produced of gradually increasing strength, until, if enough solid be used, a saturated solution is reached, at which point the reaction ceases, or equilibrium is said to be established. The concentration of this saturated solution is constant for any given temperature, but varies when the temperature is changed. When a liquid, such as water, is brought in contact with a gas, the water begins to evaporate, just as the copper sulphate began to dissolve, and finally a saturated gaseous solution is formed, above the water, which has constant concentration for a given temperature, but which increases in concentration as the temperature rises. If the temperature of the gas is lowered, the reverse change takes place, some of the water vapor condensing to liquid water, just as on cooling the saturated copper sulphate solution, some of the salt was deposited as crystals. Supersaturation is also common to both cases. If a gaseous solution, saturated with respect to water, is cooled, it may become supersaturated, but the presence of liquid, water, dust, etc., is sufficient to cause precipitation of the water.

Questions :

1. Why does the atmosphere in the neighborhood of large bodies of water contain more moisture than prairie air ?

2. Explain the occurrence of rain and fog near snow-clad mountains, over which sea air is borne by the prevailing winds.

3. Account for the formation of dew, the "sweating" of a jug of ice water, and the formation of frost on a window-pane. Why is the frost on the inside and not on the outside ?

4. Some crystals containing "water of crystallization" (see p. 75), such as Epsom salts, are generally stable in England, but often effloresce in Canada. What does this show ?

One of the influences which tend to alter the composition of the air, consists in the withdrawal of oxygen by oxidations of all kinds, *i.e.*, by rapid and slow combustions. This is counteracted by the action of green plants which give off oxygen to the surrounding air ; and the almost constant composition of the air shows that these two processes balance one another. As the processes whereby oxygen is withdrawn from the air are mainly concentrated in large towns, and as in the country a great deal of oxygen is evolved from plants, it might be expected that there would be a considerable difference in the composition of town and country air. As a matter of fact the difference is very slight indeed, the comparative constancy in composition of the air being maintained by its continual circulation.

As a result of many accurate experiments, the average composition of air, *by volume*, is found to be, approximately, 21% oxygen and 79% nitrogen, and *by weight*, 23% oxygen and 77% nitrogen. Besides the small quantities of water vapor and carbon dioxide already mentioned, air also contains small amounts of the rare gases argon, helium, neon, zenon and krypton.

Liquid Air. Air was first condensed to a liquid in 1883, by subjecting it to enormous pressure and a low temperature. It is now manufactured in large quantities at a comparatively low cost, by forcing compressed air, cooled by water, through a pipe with a small opening, into a large cylinder called a liquefier. On entering the liquefier the air expands and the temperature falls, as the expansion is accompanied by the absorption of heat. The apparatus is so arranged that the heat absorbed is taken from the compressed air on its way to the opening; thus the air in the liquefier becomes colder and colder as the process continues, until, at about − 200° C., it becomes liquid.

Liquid air has a milky appearance, owing to the presence of solid carbon dioxide, which solidifies at about − 80° C., and of ice, formed from the water vapor in the air. From the liquid air (which, of course, contains oxygen and nitrogen) the nitrogen evaporates at a lower temperature than the oxygen, its boiling point at atmospheric pressure being − 194° C., while that of oxygen is − 180° C.

Advantage is taken of this to prepare almost pure oxygen on the commercial scale, the absorption of heat on evaporation of the nitrogen being utilized to cool fresh quantities of compressed air. This liquid oxygen is transferred to cylinders, capable of withstanding great pressure, and can be withdrawn from these at any desired velocity by a properly arranged system of valves. (See p. 37.)

When liquid air is put in a tea kettle standing on a block of ice it boils vigorously, as the ice is very hot in comparison with the temperature of the liquid air in the tea kettle. If water is poured into a kettle containing liquid air, the liquid air boils over as the water is so much hotter, while the latter instantly freezes, as the liquid air in vaporizing uses up heat. If the kettle containing liquid air is placed over the flame of a spirit lamp, frost and ice collect on the outer surface, the intense cold of the kettle solidifying the carbon dioxide and the water vapor, which are the products of combustion of wood alcohol; the moisture in the air surrounding the kettle is also changed to ice. By the use of liquid air mercury may be frozen so hard that it may be used as a hammer, while an iron or tin vessel becomes so brittle that it may be readily rubbed to a powder.

APPLICATIONS OF THE PRINCIPLES OF CLASSIFICATION.

Liquid air and foggy air are good examples to illustrate the application of the classifications explained on page 114.

"*Liquid air*" is a *mechanical mixture* of a solid chemical compound, carbon dioxide (causing the milkiness referred to above), with a liquid solution. The lack of homogeneity evidenced by the milkiness is in this case sufficient to show that a mechanical mixture is present. The solid carbon dioxide would be found to have the same composition, whatever the composition of the air might be, from which it was condensed; for example, if the percentage of oxygen had been increased by adding more oxygen, the composition of the solidified carbon dioxide would not have been altered. Consequently it is a *chemical compound*. The part of the "liquid air" actually liquid "contains" oxygen and nitrogen, but since the composition of this liquid is found to vary gradually during the process of preparation, and indeed would also be different if produced from an air containing, say, more oxygen, it is not a chemical compound of oxygen and nitrogen. Further, since its properties are *not* the individual properties of liquid oxygen and liquid nitrogen (vapor tension, for instance), it is *not* a mechanical mixture, but a liquid *solution*.

Foggy air is a *mechanical mixture* of liquid water (in fine drops) with a gaseous solution. The gaseous solution contains oxygen, nitrogen, carbon dioxide, water vapor, and a few other gases.

PROBLEMS.

1. One gram of magnesium is completely burned in 10 litres of air, measured at 17° C. and 760 mm. pressure. Find the weight of magnesium oxide, MgO, formed, and the volume of gas left, measured at the same pressure and temperature.

2. 100 litres of air at 760 mm. and 20° C. are passed over a weighed quantity of lime. The lime is then found to have increased in weight by 0.025 grams, owing to the absorption of the carbon dioxide from the air. Assuming that the whole of the carbon dioxide has been removed from the air by this process, calculate the percentage by volume of carbon dioxide in the air.

3. Air was drawn through a tube containing lumps of calcium chloride, which removed all the water vapor by forming a hydrated salt. The increase in weight of the calcium chloride tube, after passing 10 litres of air at 750 mm. and 17° C., was 0.080 grams. Find the per cent., by volume, of water vapor in the air.

4. 75 cc. of a gas containing only oxygen and nitrogen, in unknown proportions, are shaken with a solution of pyrogallic acid and potash to absorb the oxygen. After restoring the original temperature and pressure, the remaining volume is found to be 30 cc. Find the per cent. composition by volume, and the per cent. composition by weight, of the gas. Formula of nitrogen is N_2.

5. Sufficient phosphorus to remove all the oxygen was burned in a closed tube containing 100 cc. of air at 27 C. and 740 mm. Find the pressure of the remaining gas at 37° C.

6. 1000 cc. of carbon dioxide at 760 mm. pressure and 15° C. are shaken for a short time with 100 cc. of water. The pressure of the gas is then found to be reduced to 700 mm., the temperature being the same. Find the weight of carbon dioxide (CO_2) dissolved in the water.

7. If the composition of normal air by weight is 76.995 per cent. of nitrogen and 23.005 per cent. of oxygen, what is the per cent. of each constituent by volume? Formula of nitrogen is N_2.

8. A litre of dry air at 0° and 760 mm. is passed over heated copper. The copper increased in weight 0.297 g. What per cent. of oxygen by weight is contained in this sample of air?

9. What is the difference in weight between 10 l. of air at 15° C., and 10 l. of air at 775.62 mm. (22.4 l. of air at 0° C. and 760 mm. weigh 28.97 g.).

CHAPTER XIV.—NITROGEN AND ITS COMPOUNDS

NITROGEN.

It has been shown that there is present in the air, besides oxygen, another substance which constitutes the greater part, both by weight and by volume. This substance is nitrogen. It is a colorless, odorless gas, only slightly soluble in water. Its reacting weight, N, is 14.04 g.; 22.4 litres at 0° C. and 760 mm. weigh 28.08 grams. Its formula weight is, therefore, 28.08 as represented by the formula N_2. Nitrogen is distinguished from oxygen and from most other gases by the fact that it is neither combustible nor supports combustion, and has little tendency to take part in chemical reactions.

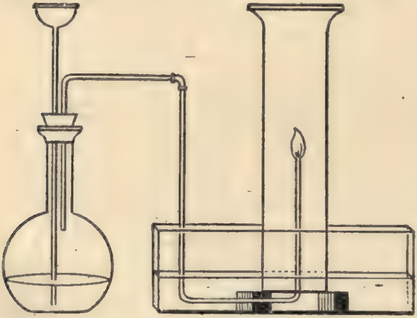

Fig. 37.

In order to prepare nitrogen, it is only necessary to remove the oxygen from the air by a substance that forms a solid, or liquid, soluble oxide. This may be done by means of phosphorus, as in experiment 1, p. 155, or by burning a jet of hydrogen in a vessel over water, as in Fig. 37. Hydrogen is generated in the flask, from zinc and hydrochloric, or dilute sulphuric acid. *After all air has been expelled*, ignite the hydrogen and bring down over the jet a glass cylinder. As soon as the hydrogen ceases to burn, disconnect the delivery tube. (Why?)·

The oxygen may also be removed by slowly passing air over strongly heated copper turnings in a glass tube. The reaction is often represented by the equation—

$$2\,Cu + O_2 = 2\,CuO.$$

This is another example of the loose way in which equations are used by chemists. The reaction takes place between copper and a gaseous solution, giving copper oxide and either nitrogen or a gaseous solution containing a smaller proportion of oxygen than that present in air. The equation, therefore, does not really represent what takes place, but for reasons similar to those given on page 145, it is often found convenient to use chemical formulae in this way.

Nitrogen may be prepared also by heating ammonium nitrite, a substance whose composition is represented by the formula NH_4NO_2. When this is heated, a reaction takes place expressed by the following equation :—

$$NH_4NO_2 = 2\,H_2O + N_2.$$

Nitrogen is essential to the growth of plants, and is absorbed by them, chiefly in the form of compounds dissolved in the water and taken in through the roots.

The roots of many plants belonging to the clover family, such as peas, beans, clovers, vetches, etc., serve as hosts for certain bacteria that have the power to form nitrates from the nitrogen and oxygen of the air. These nitrates dissolve in the water in the soil, and are absorbed by the roots of succeeding crops.

Nitrogen compounds, particularly nitrates, are present in large quantities in all barnyard manure, and they constitute some of the most important ingredients of all fertilizers. During thunderstorms, the lightning, passing through the moist air, causes the nitrogen to form compounds with oxygen; these dissolve in the rain, are carried to the earth and are there of service to plant life.

Quite recently, compounds of nitrogen, of value as fertilizers, have been produced by the discharge of electricity through air. The commercial application of this process has been attempted at Niagara Falls.

COMPOUNDS OF NITROGEN.

Nitrous Oxide, Formula N_2O.

SUPPLIES: Ammonium nitrate, large test tube, cork and delivery tube, litmus paper, splinters of wood, bottles for collecting gas, sulphur, phosphorus, magnesium.

Preparation. Into a large test tube put ammonium nitrate to the depth of two inches. Connect the test tube, by means of a cork and delivery tube, to the mouth of a bottle filled with water and inverted in a pneumatic trough. Heat the test tube and collect several bottles of the gas that comes off.

The proportions of the substances involved are represented by the equation—

$$NH_4NO_3 = 2\ H_2O + N_2O.$$

Experiments with Nitrous Oxide:

1. Into one of the bottles filled with the gas thrust a glowing splinter. If no reaction occurs, use a lighted splinter.

What properties of nitrous oxide are illustrated by this experiment?

The combustion of other substances, such as sulphur, phosphorus, or magnesium may also be tried in jars of the gas.

2. Mix together some hydrogen and nitrous oxide, and ignite them.

Is there any evidence of the formation of water?

3. Place another bottle mouth downwards in a vessel partly filled with water and shake it gently, being careful to keep the mouth of the bottle below the surface of the water. Test the water with blue litmus paper.

What properties of the gas are shown by these experiments? How may the gas be distinguished from oxygen? What happens to a portion of the gas when it is collected over water? The gas is sometimes collected over warm water. Why? (See p. 100.) If the formula of the gas is N_2O, what is the ratio of its density to that of hydrogen?

The density of a gas is sometimes expressed, not as the weight of 22.4 litres at 0° C. and 760 mm., but as the ratio of the weight of any volume of the gas to that of an equal volume of hydrogen, or of air, at the same temperature and pressure. An advantage of this method of representation is that the numbers do not vary with the temperature and pressure.

Nitrous oxide is a gas containing two reacting weights of nitrogen to one of oxygen. It is colorless, has a faint sweetish odor, and, at the temperature of the laboratory, is soluble in its own volume of water. When dissolved in water, the solution is said to contain a weak acid, known as hyponitrous acid, with the formula HNO. Note that $H_2O + N_2O$ adds up to 2 HNO. This substance has recently been prepared and is very explosive; a few of its salts have been known for a long time. Even before the preparation of hyponitrous acid, however, the name and formula were used. (See p. 149.)

Nitrous oxide decomposes, on being heated, into nitrogen and oxygen, and if too much heat be applied in its preparation from ammonium nitrate, the gas collected will contain free nitrogen and oxygen. Many substances that will burn in air will also burn in nitrous oxide.

Nitrous oxide, when inhaled, forms a compound with the blood and produces unconsciousness and insensibility to pain. It is therefore used as an anaesthetic and was formerly known as "laughing gas," because of the peculiar effect on persons inhaling it. The gas cannot be decomposed in the body in such a way that its oxygen becomes available to support life, therefore, when it is to be inhaled for a lengthened period, oxygen must be mixed with it, in the same proportion as the former is found in air.

Nitric Oxide, Formula NO

SUPPLIES: Copper turnings, nitric acid, flask, or large test tube, cork and delivery tube, bottles for collecting gas, solution of ferrous sulphate, splinters of wood, phosphorus, deflagrating spoon.

Preparation. Place some copper clippings in a flask or large test tube and pour on these some nitric acid. Fit the flask with a cork and delivery tube, heat gently, and collect the gas over water.

Note the color of the gas in the flask and in the bottles.

Experiments with Nitric Oxide:

1. Raise one of the bottles filled with the gas out of water, and after leaving it uncovered for a short time replace it mouth downwards.

What was observed when the gas was in contact with air, and when the bottle was replaced in water?

2. Introduce into a bottle filled with the gas, (*a*) a glowing splinter, (*b*) a burning splinter, (*c*) a small piece of brightly burning phosphorus.

3. Place a bottle filled with the gas, mouth downwards in a vessel partly filled with a solution of ferrous sulphate.

Nitric oxide is a colorless gas which is insoluble in water. 1 litre of the gas weighs 1.341 grams at 0°C. and 760 mm., therefore, 22.4 l. of the gas would weigh 30.04 grams, which is the formula weight represented by NO.

When nitric oxide is exposed to air, it combines with oxygen and forms the compound nitrogen peroxide, NO_2, which is a brown gas. If pure oxygen be passed into nitric oxide standing over water, the volume will decrease instead of increasing, as the compound formed is readily soluble in water. By careful addition of oxygen all the gas in the bottle may be made to disappear.

Nitric oxide readily dissolves in a solution of any ferrous salt, giving the liquid a dark brown color. This, together with the formation of brown fumes when it escapes into air, forms a convenient *test* for the gas.

If heated sufficiently, nitric oxide may be decomposed into nitrogen and oxygen. For this reason brightly burning phosphorus continues to burn when placed in the gas, while a burning splinter is extinguished.

The two oxides of nitrogen, nitrous oxide and nitric oxide, give a good illustration of another convention used in chemical nomenclature. When an element combines with oxygen in more than one proportion, the two oxides formed are distinguished by adding the terminations *-ous* and *-ic* to the root of the name of the element; the first ending being used for the compound containing the lesser proportion of oxygen, and the second for the one with the larger proportion. Other examples are ferrous oxide, FeO, and ferric oxide, Fe_2O_3; mercurous oxide, Hg_2O, and mercuric oxide, HgO. The system is extended also to other compounds besides the oxides, for example:—

Ferrous chloride, $FeCl_2$, and ferric chloride, $FeCl_3$.
Phosphorus chloride, PCl_3, and phosphoric chloride, PCl_5.
Ferrous sulphate, $FeSO_4$, and ferric sulphate, $Fe_2(SO_4)_3$.

In the latter case it is the difference in the proportion of the iron to the substance represented by the group of symbols, SO_4, that is indicated by this nomenclature.

Nitrogen Peroxide, Formula NO_2.

Nitrogen peroxide may be formed by heating lead nitrate, which decomposes as follows:—

$$2\ Pb(NO_3)_2 = 2\ PbO + 4\ NO_2 + O_2.$$

The peroxide is a brown gas which dissolves in water, yielding an acid solution.

The density of the gas at ordinary temperature, taking that of hydrogen as 1 (see p. 169), is 44. This gives the formula weight as 89, which corresponds *approximately* to the formula N_2O_4. At about 100° C., the density as compared with hydrogen, is 22.1; therefore, at this temperature the formula weight is 46, represented *approximately* by the formula NO_2.

To express these facts, different formulae must be assigned to the gas at these two temperatures. Between these two temperatures it is found that the density varies gradually between the two limits represented by the formulae N_2O_4 and NO_2. All these facts may be satisfactorily accounted for by the assumption that, at the intermediate temperatures, the gas is a solution of two substances, one having the formula N_2O_4, and the other NO_2, and that as the temperature is raised, more and more of the latter is formed at the expense of the former.

Other Oxides of Nitrogen.

There is another oxide of nitrogen called nitrogen pentoxide, or nitric anhydride, formula N_2O_5, which dissolves in water, forming nitric acid.

A substance called nitrogen trioxide has also been described, but there is now some doubt as to its existence as a chemical compound.

NITRIC ACID, HNO_3.

SUPPLIES: Glass retort, flask, retort stand, potassium or sodium nitrate, copper turnings, evaporating dish, splinter, solutions of indigo and potassium permanganate, litmus paper, a carbonate, caustic potash.

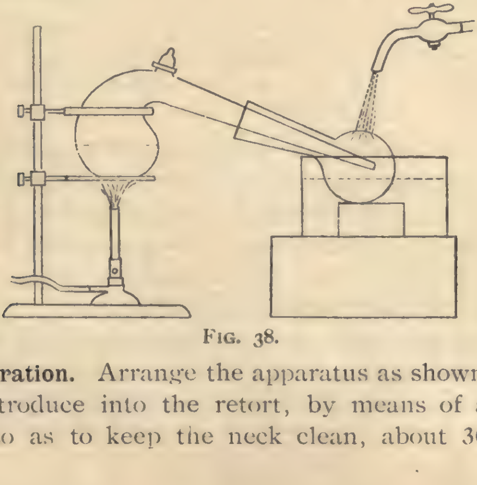

FIG. 38.

Preparation. Arrange the apparatus as shown in Fig. 38. Introduce into the retort, by means of a folded paper so as to keep the neck clean, about 30 grams of potassium or sodium nitrate, and pour upon it through a funnel, about the same weight of concentrated sulphuric acid. Heat the retort and allow cold water from the tap to pour gently over the flask.

Experiments with Nitric Acid:

Dilute some of the liquid obtained with fifteen or twenty times its volume of water, and with this solution make the following experiments:—

1. Taste the liquid *carefully*, and test with litmus.

2. To some of the liquid add a little baking soda, or any carbonate.

To what class of substances does the solution belong? Why?

3. Neutralize some of the solution with potassium hydroxide and evaporate to dryness in a dish. Taste a little of the solid obtained. Scrape as much as possible of the substance from the dish and heat it in a test tube over a flame. When melted and bubbling freely, add a small piece of charcoal. After the reaction has ceased, test the gaseous contents by inserting a glass rod wet with lime water. What happened to the carbon?

The solid obtained in experiment 3 was nitre, or potassium nitrate, as may be proved by a comparison of its properties with those of a specimen of that substance.

Nitre, whose composition is represented by KNO_3, is one of a large class of substances including Chili saltpetre, $NaNO_3$, lead nitrate, $Pb(NO_3)_2$, etc., all of which can be prepared by the evaporation of solutions made by neutralizing various bases by nitric acid. They are therefore salts, and are given the common name *nitrates*.

Formula of Nitric Acid.

As already explained in a previous chapter, an acid solution is often said to contain a chemical compound with a definite formula, although no such chemical substance may be actually known. Now nitric acid may be made by dissolving nitrogen pentoxide, N_2O_5, in water, and we might therefore say that nitric acid solution contains N_2O_5, or $N_2O_5 + H_2O$, or $N_2O_5 + 2\,H_2O$, etc. The second of these, namely $N_2O_5 + H_2O$, has been selected as being the simplest of them, which contains the H necessary to suggest the acid properties. (See p. 149.) Another advantage of this formula (which is also the probable historical reason for its use), is that it represents very closely the composition of the strongest acid obtained by the ordinary method of preparation. The composition represented by the formula $N_2O_5 + H_2O$, is represented equally well by either $H_2N_2O_6$ or HNO_3, and as the latter is the simpler of the two, this is now taken as the formula of the substance that nitric acid is said to contain. In ordinary chemical language it is often said to be the "formula of nitric acid."

Valency of Radicals.

On looking at the formulae of the nitrates recorded above, it will be seen that combined with one reacting weight of the monovalent metals there are amounts of oxygen and nitrogen represented by the group of symbols NO_3, while with one reacting weight of the divalent metals, the amounts represented by $2(NO_3)$, indicated in formulae by $(NO_3)_2$, are combined. Thus we can extend the idea of valency, and say that the oxygen and nitrogen together, represented by NO_3, are monovalent. Consequently the number of H's in an acid formula indicates the valency of the rest of the formula, or, as it is often expressed, the valency of the "acid radical." Thus the H of an acid formula may be replaced by the symbol of a monovalent metal to give the formula of a salt, for instance HNO_3, nitric acid, and KNO_3, potassium nitrate. In the case of a divalent metal, like calcium, the formula of the salts will contain two monovalent acid radicals, *e.g.*, $Ca(NO_3)_2$. The formula H_2CO_3 (p. 149) was selected for the hypothetical hydrogen carbonate, spoken of as being contained in carbonic acid, to correspond with the formulae of its salts, Na_2CO_3, K_2CO_3, etc. The acid radical CO_3 is, therefore, divalent. The terms *monobasic acid* and *dibasic acid* are used for acids containing a monovalent and divalent acid radical respectively. A knowledge of the valency and basicity is thus a great help in remembering the formulae of salts.

SUPPLIES: Nitric acid, zinc, copper, magnesium, iron, charcoal, sulphur, solutions of indigo and ferrous sulphate.

Experiments with Nitric Acid and Nitrates:

1. Into a test tube containing some strong nitric acid, drop a small piece of zinc. Apply a lighted match to the mouth of the test tube.

Is the gas evolved hydrogen? Has it the appearance of hydrogen?

2. Repeat experiment 1, using pieces of copper, magnesium and iron instead of the zinc.

How does nitric acid differ from other acids that you are familiar with?

3. Fuse some nitre in a test tube and drop in a small piece of charcoal. Repeat, but drop in a piece of sulphur.

What gas appears to be formed?

4. Heat some strong nitric acid in an evaporating dish until it boils, then hold a piece of glowing charcoal close to the surface of the liquid.

5. Drop some nitric acid on a piece of black cloth.

What property of nitric acid and nitrates is illustrated by these experiments?

Metals like zinc, magnesium, iron, etc., which react with hydrochloric and sulphuric acids, forming chlorides and sulphates, and liberating hydrogen, also react with nitric acid, giving nitrates. But since nitric acid yields its oxygen quite readily (or is an *oxidizing* agent), the hydrogen replaced by the metal is oxidized by some of the nitric acid, forming water and a mixture of various oxides of nitrogen. This accounts for the fact that a much greater quantity of hydrogen is liberated when these metals react with hydrochloric or sulphuric acid, than when nitric acid is used. Indeed, when strong nitric acid is used, no hydrogen at all is liberated.

The salts of nitric acid are also powerful oxidizing agents, as is illustrated by Exp. 3, where the fused nitre gave up oxygen to the carbon, changing it into carbon dioxide, and to the sulphur, forming sulphur dioxide. The gases evolved in experiments 1 and 2 contain various oxides of nitrogen—the "reduction products" of nitric acid. The term *reduction* is used for processes in which oxygen is removed from a substance; and materials which have the property of depriving other substances of part or all of their oxygen are called *reducing agents*. Thus, in the above experiments, the metals, and carbon and sulphur were reducing agents, since they deprived the nitric acid of part of its oxygen, leaving other nitrogen compounds.

It must be noted, however, that the terms *oxidation* and *reduction* are also used in a wider sense. Any process involving the increase in valency of a metal is spoken of as oxidation, while for a change involving a decrease in valency the term *reduction* is used. For example, in the change of ferrous chloride ($FeCl_2$) into ferric chloride ($FeCl_3$), which may be brought about by chlorine, nitric acid, or potassium permanganate, the valency of iron increases from 2 to 3, and the salt is said to be "oxidized," although it contains no oxygen at all. Similarly, when mercuric chloride ($HgCl_2$) is changed to mercurous chloride ($HgCl$) by chloride of tin (stannous chloride), the process is called "reduction," and the chloride of tin is said to "reduce" the mercuric salt. Thus ferrous chloride and stannous chloride are "reducing agents," and nitric acid, chlorine, and potassium permanganate are "oxidizing agents."

Test for Nitrates. To about 1 cc. of a solution of any nitrate, add a small quantity of a solution of ferrous sulphate, and pour gently down the side of the tube some strong sulphuric acid. A brown ring or coloration will be observed at the boundary between the two liquids. Some of the ferrous sulphate is "oxidized" to ferric sulphate by the nitric acid, and the latter is reduced to oxides of nitrogen, one of which gives the brown color with the remaining ferrous salt. (See p. 171.)

Notes on Nitric Acid and Nitrates.

The most concentrated nitric acid is a colorless liquid, with a specific gravity of 1.53 and boiling point 86° C. It does not keep well, for even under the influence of light it decomposes into oxygen and compounds of nitrogen, chiefly the oxides, which dissolve in the rest of the acid, giving it a yellowish color. Addition of water makes it more stable. The sp. gr. of the commercial acid is about 1.42, and it contains from 60 to 70 per cent. of the acid represented by HNO_3, the rest being water.

Large quantities of nitric acid are used in the manufacture of nitrates, dyestuffs, sulphuric acid, nitroglycerine, guncotton, in the refining of gold and silver and in etching copper plates.

Industrial Preparation of Nitric Acid. Practically all nitric acid is now made by heating sodium nitrate (Chili saltpetre) with sulphuric acid in cast iron retorts.

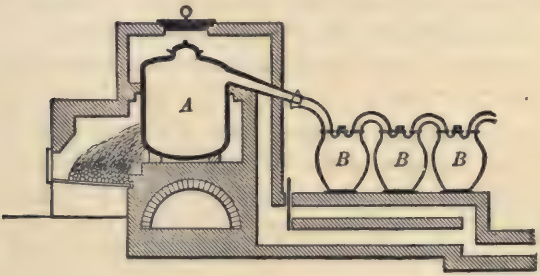

FIG. 39.

Fig. 39 shows one form of plant, although there are many modifications of the process, principally improved methods of condensing the distilled acid vapor. The retort A is entirely surrounded by the flames from the grate. The acid vapor is condensed in the series of large glass or earthenware bottles (B B). The last bottle in the series leads into a tower filled with coke, over which water trickles to condense the escaping fumes. The first two or three bottles contain the most concentrated acid, but it is contaminated with oxides of nitrogen and sulphuric acid. Owing to the decomposition of some of the nitric acid by heat, the condensed acid has a yellow or red color, due to the presence of other oxides of nitrogen. The acid is "bleached" by heating it to boiling point and blowing in warm air (Thorp's "Outlines of Industrial Chemistry").

The salts of nitric acid, particularly potassium nitrate (saltpetre) and sodium nitrate (Chili saltpetre), occur in large quantities in various countries. Wherever decomposition of organic matter, particularly animal excrements, takes place in a soil containing basic substances, and subject to the action of the oxygen of the air, nitrates are formed. This process is of the greatest importance to the human race, as through it the useless nitrogenous substances of dead plants and animals are converted back into the nitrates which are essential to the growth of living plants.

Nitrates are also formed from the basic constituents of the soil together with the oxygen and nitrogen of the atmosphere. This reaction is brought about by the influence of "nitrifying bacteria" in the roots of certain plants referred to on p. 167, and by electrical discharge in the atmosphere, the products of which are washed down by the rain.

All nitrates are soluble in water and decompose on being heated, usually in the same way as lead nitrate.

$$2\ Pb(NO_3)_2 = 2\ PbO + 4\ NO_2 + O_2.$$

The exceptions to this method of decomposition are ammonium nitrate and potassium nitrate.

Ammonium nitrate, when heated gently, decomposes into water vapor and nitrous oxide, as described on page 168.

Potassium nitrate, on being heated, decomposes into potassium nitrite and oxygen, as represented by the equation—

$$2\ KNO_3 = 2\ KNO_2 + O_2.$$

The substance, potassium nitrite, obtained by this last reaction, is one of a series of salts called nitrites, corresponding to nitrous acid, HNO_2. (For nomenclature, see chlorates and chlorites, p. 150.) Nitrites are generally prepared by the reduction of nitrates.

Nitrites are also formed when organic matter decomposes, and as decomposition is caused by bacteria, the presence of nitrites in water usually indicates the presence of bacteria, some of which cause disease.

All nitrites may be oxidized and converted into nitrates. For a similar reason nitrous acid cannot be kept for a length of time, as it absorbs oxygen and is converted into nitric acid.

AMMONIA, NH_3.

SUPPLIES: Liquor ammoniæ (solution of ammonia in water), flask, or large test tube, cork and glass tube, bottles, litmus paper, hydrochloric acid, evaporating dish.

Preparation. Pour a small quantity of ammonia solution into a flask, or test tube, fit with a cork and glass tube. Heat the flask and hold a bottle, mouth downward, over the open glass tube (Fig. 40).

What common property of solutions of gases in liquids is illustrated by this experiment?

FIG. 40.

Experiments with the gas obtained:

1. Carefully smell the gas by wafting it gently towards you.

2. Thrust a lighted splinter into a jar filled with the gas.

3. Place a thoroughly dried bottle, filled with the gas, mouth downwards in water.

4. Bring near the end of the delivery tube (*a*) a piece of moistened red litmus paper, (*b*) a glass rod moistened with hydrochloric acid.

5. Place two evaporating dishes, one containing ammonia solution, and the other hydrochloric acid, under a bell jar, and allow them to stand until next day.

Make a list of all the properties of the gas as shown by the above experiments.

The gas used above, obtained by heating the solution in water, is called ammonia. Analysis shows that it contains 17.75% of hydrogen and 82.25% of nitrogen. To find the proportions in terms of reacting weights, these amounts must be divided by the reacting weights of hydrogen and nitrogen respectively, thus :—

17.75 grams of hydrogen to 82.25 grams of nitrogen represent the same proportion as

$$\frac{17.75}{1.01}\text{ R.W. of hydrogen to }\frac{82.75}{14.04}\text{ R.W. of nitrogen;}$$

or, 17.58 H to 5.860 N.

By the law of reacting weights, these numbers must be in the proportion of small round numbers ; and by dividing through by the smaller, we get

3 H to 1 N.

The formula of ammonia must therefore be NH_3, or N_2H_6, or N_3H_9, etc. But since at 0° C. and 760 mm., 22.4 litres of the gas are found to weigh 17.07 g., the formula is NH_3.

Ammonia Solution.

SUPPLIES: Litmus paper, linseed oil, or sweet oil, ammonium hydroxide, solutions of potassium, sodium, or calcium hydroxides.

Experiments:

Make the following tests with a solution of ammonia:—

1. Test with litmus.

2. Dilute the solution, and taste cautiously.

3. Moisten the thumb and forefinger with the solution, and rub them together.

4. Shake some of the solution with sweet oil, or linseed oil.

5. Repeat all these experiments, using (*a*) potassium hydroxide solution, (*b*) sodium hydroxide solution, (*c*) lime water (calcium hydroxide solution).

It will be seen from the above experiments that all these solutions have several common properties. They are therefore put in one class, and called bases or alkalies. (The terms base and alkali are also used commonly for chemical substances which, when dissolved in water, give solutions which are bases. Compare this with the use of "carbonic acid" for carbon dioxide, etc.)

The formulae of the compounds, solutions of which were used in the above experiments, are as follows: Potassium hydroxide, KOH, sodium hydroxide, $NaOH$, calcium hydroxide, $Ca\ (OH)_2$. It will be noticed that there is a similarity in the formulae, all of them containing OH. Now when ammonia, NH_3, is dissolved in water, an alkaline solution is formed, and it will therefore be convenient to look upon such a solution as if it "contained" a compound with a formula analogous to those of potassium hydroxide and the other hydroxides. This can be done by saying that the solution of ammonia in water contains ($NH_3 + H_2O$), which may be written NH_4OH, thus suggesting the basic properties of the solution. It will be remembered that a very similar convention was used in writing the formulae of acids (p. 149).

It is convenient also to have a name for the compound NH_4OH, although no such compound has ever been made. It is called ammonium hydroxide, the word "hydroxide" indicating the similarity to the other bases, while the word "ammonium" (used for the set of symbols NH_4) suggests its preparation from ammonia. As will be seen later, this term is very convenient in naming the salts which can be formed by neutralizing the solution of ammonium hydroxide by acids. The set of symbols OH, which occurs in the formulae of all hydroxides, has likewise received a name, viz., "hydroxyl." No substance "ammonium" with the formula NH_4 has ever been prepared, and the same is true of "hydroxyl." It is, nevertheless, customary to speak of bases as "containing" hydroxyl, and of ammonium compounds as "containing" ammonium. The formulae of compounds both of hydroxyl and of ammonium may be remembered if we assign the constant valency *one* to both radicals.

The chief ammonium salts are ammonium chloride, ammonium sulphate, ammonium nitrate, and ammonium carbonate.

Ammonium Chloride, or Sal Ammoniac, NH_4Cl.

Experiments:

1. Fill two dry bottles, one with hydrogen chloride and the other with ammonia gas. Cover with glass plates, invert the hydrogen chloride bottle over the other and slip out the plates. (Compare Exps. 4 (*b*) and 5, p. 182.)

2. In the bottom of a test tube, heat strongly a small crystal of dry ammonium choride, keeping the upper part of the test tube as cool as possible. Compare the taste of the substance condensed on the side of the test tube with that of the original material.

In experiment 1 the two gases reacted to form a solid substance with a composition represented by NH_4Cl. It is ammonium chloride and is identical with the salt that might be obtained by neutralizing a solution of ammonia with hydrochloric acid and evaporating the water. The equation for the reaction is

$$NH_3 + HCl = NH_4Cl.$$

The second experiment is an example of sublimation, a term applied to the change of a solid to a gas and back again without melting. Ice and many other solids evaporate slowly below their melting points, and may be sublimed if proper precautions be taken, but in the case of ammonium chloride, the pressure with which the gas is given off becomes greater than atmospheric pressure before the melting point is reached, and sublimation can be, therefore, easily effected. Iodine afforded another example of the same phenomenon in experiment 7, p. 11. The process of sublimation may be used for purifying such solids, just as distillation is used for purifying water.

The density of the gas obtained by subliming ammonium chloride does not correspond to the formula NH_4Cl, but is close to that calculated for a gaseous solution of equal volumes of ammonia and hydrogen chloride. We consequently say that ammonium chloride decomposes, when heated, into these two gases according to the equation

$$NH_4Cl = NH_3 + HCl,$$

and that the reverse takes place on cooling.

Ammonium Sulphate $(NH_4)_2SO_4$.

Experiment :

Mix a little ammonium sulphate and slaked lime, place it in a large test tube and heat. Smell the gas that comes off. Test it with litmus. What is it?

Ammonium chloride may be substituted for the sulphate in this experiment.

Ammonium sulphate is a salt produced in large quantities as a by-product of the gas works. It is extensively used as a fertilizer on account of the nitrogen which it contains. It is also used for the preparation of ammonia, the reaction taking place being represented by the equation

$$(NH_4)_2SO_4 + Ca(OH)_2 = CaSO_4 + 2\ NH_3 + 2\ H_2O.$$

Tests for Ammonia and Ammonia Solution. Since even the weakest solutions of ammonia give off ammonia gas, the solution may be recognized by the characteristic odor, by the formation of white clouds of solid ammonium chloride with hydrogen chloride, or by the change in color of moistened red litmus paper held above it.

1. "Nessler's reagent" for solutions of ammonia may be made as follows :—

To a solution of mercuric chloride add a solution of potassium iodide, until the precipitate formed just ceases to be redissolved; then add an equal volume of a strong solution of caustic potash. The solution so prepared gives a reddish precipitate with even slight traces of ammonia.

2. Test for ammonium salts: To a small quantity of any ammonium salt in a test tube add a solution of potassium or sodium hydroxide, and heat. Smell the gas that comes off.

Notes on Ammonia.

Ammonia is a colorless gas having a very pungent odor. Its density as compared with hydrogen is 8.45. At the temperature of the laboratory, one litre of water will dissolve about 800 litres of the gas. This solution is known commercially as liquor ammoniæ, or spirits of hartshorn, a name due to the fact that ammonia was formerly prepared by the distillation of horns, hoofs, hair, etc. The gas can be completely removed from the solution by boiling.

On account of its alkaline properties, ammonia solution is used as a cleaning agent, especially for the removal of grease. It is also used in dyeing, in calico printing, in the manufacture of dyestuffs, soda, etc., and as a restorative in cases of fainting.

Ammonia may readily be converted to a liquid by pressure or cold. When the pressure is released the liquid vaporizes, and in doing so uses up a great quantity of heat, thus lowering the temperature of the surroundings. These facts are taken advantage of in the artificial manufacture of ice, and in refrigerating plants in factories, meat packing houses, storehouses, ships, hotels, etc.

Fig. 41 is a diagrammatic representation of an ammonia refrigerator, such as is almost exclusively used at present. P is a pump which exhausts the gas from the pipe G, and forces it through C into the condenser K. The gas, under pressure in the coils of the condenser, becomes liquefied, and is cooled by a stream of cold water entering at A and escaping at B. The liquid ammonia escapes through the regulating valve R, into the evaporating vessel V, where it is gasified, because of the low pressure caused by the exhaust action of the pump through the pipe G.

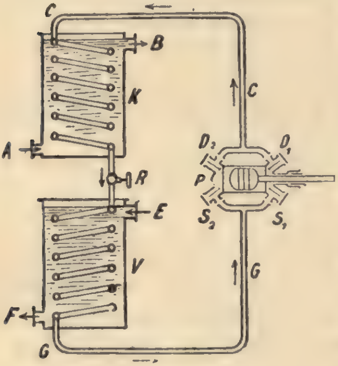

FIG. 41.

The salt solution that circulates through the pipes, with which the apartment to be cooled is fitted, enters the evaporating vessel at E and passes out at F. Heat is absorbed from the salt solution as it passes through this vessel, by the vaporizing ammonia, and it is cooled to any required temperature by regulating the flow of the brine and the liquid ammonia by means of the regulating valve R. (Ostwald's "Principles of Chemistry.")

Industrial Preparation. Ammonia is prepared from ammonium salts, which are obtained in large quantities as by-products in the manufacture of illuminating gas from the distillation of soft coal (see page 111). These salts are heated with a base, usually slaked lime; the ammonia gas generated is passed into water. The aqueous solution thus obtained is sold as liquor ammoniæ.

CHAPTER XV.—SULPHUR AND ITS COMPOUNDS

SULPHUR.

SUPPLIES: Roll sulphur, flowers of sulphur, carbon disulphide, Hessian crucible, or evaporating dish, test tube, mercury, mortar and pestle.

Experiments with Sulphur:

1. Examine a piece of sulphur and describe its properties. Find its specific gravity. (See Exp. 2, p. 8.)

2. Place a small piece of roll sulphur in a test tube, add distilled water and shake. Heat the water.

Does water dissolve sulphur?

Evaporate a drop of the liquid to dryness on a watch glass. Is there a residue?

3. Pour off the water, and to the sulphur add carbon disulphide, and shake. Does the sulphur dissolve?

Pour the liquid contents of the test tube into an evaporating dish and set it away until next day. Examine carefully what is left.

4. Repeat the above experiment, using flowers of sulphur.

5. Fill a Hessian crucible, or evaporating dish, with flowers of sulphur, support it on the ring of a retort stand and heat until it melts. Allow it to cool, and when a crust has formed, break two holes in it, pour out the liquid sulphur and set the vessel away to cool.

What is seen projecting from the sides of the vessel? Are these crystals the same shape as those formed in experiment 3?

6. About half fill a test tube with flowers of sulphur and heat until the sulphur begins to boil. Note carefully all changes that take place in the sulphur. What collects on the upper portion of the tube, or escapes from the mouth?

When the sulphur begins to boil, pour the liquid into cold water. Describe the mass of sulphur obtained in this way. Keep a portion of it for several days and note any changes.

7. Place a small piece of this plastic sulphur in a test tube, add carbon disulphide and shake. Is it soluble?

8. Place some roll sulphur, flowers of sulphur or plastic sulphur, in a deflagrating spoon; ignite, and lower it into a bottle containing air. Carefully smell the fumes.

Add some water, place the hand tightly over the mouth of the bottle and shake. Test with litmus.

9. Rub a piece of sulphur on a bright silver or copper coin.

10. Put a drop or two of mercury and a little flowers of sulphur in a mortar and rub for some time with the pestle.

11. Powder some iron pyrites, FeS_2; put it in a hard glass tube and heat strongly.

What collects on the cool sides of the tube?

If free access of air had been allowed, what would probably have been formed? Try the experiment in a tube open at both ends.

Notes on Sulphur.

Most of the sulphur used in the industries is derived from the native sulphur which is found in volcanic regions, mixed with clay, finely divided rock, and other matter thrown out by the volcano.

By far the largest part of the world's supply of sulphur comes from Sicily, but some is obtained in Italy, Greece, Japan, and the United States. The annual output of Sicily is about 500,000 tons; and on account of the favorable position of this island as a shipping point, the abundant supply of cheap labor, and the extensive deposits, this will probably continue to be the principal source of supply.

The sulphur left after the evaporation of a carbon bisulphide solution is in the form of *orthorhombic* crystals, see Fig. 42, while that obtained by melting sulphur and allowing it to cool, is in long needle shaped, *monoclinic*, crystals, see Fig. 43. Since the monoclinic crystals may be converted *weight for weight* into the rhombic; in other words, since the monoclinic crystals may be converted into the rhombic, without the addition of any other substance, and without any other material being formed, the two kinds of crystals must have the same composition. Sulphur being an element, these are therefore called *allotropic modifications of sulphur*. (See

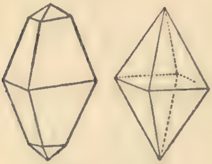

FIG. 42.

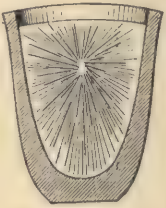

FIG. 43.

pp. 73 and 91.) Both of these forms of sulphur combine with oxygen, giving the same gas—sulphur dioxide; but this fact alone is not sufficient to show that the two kinds of crystals are allotropic modifications of the same element, since iron pyrites and many other substances also combine with oxygen, giving sulphur dioxide. In order to show the existence of allotropic modifications by this method, it would be necessary to show that *the same weight of the two substances in question gave products identical in kind and weight.*

Several other modifications of sulphur are known, for instance, the plastic sulphur obtained in experiment 6, and the sulphur suspended in what is known as "milk of sulphur," produced by adding an acid to a solution of sodium thiosulphate (hyposulphite of soda).

When sulphur is heated to about 120° C. it melts to a thin amber colored liquid. As the temperature is raised the liquid darkens and becomes viscous, until at about 250° C. it is almost black and too thick to pour. On being heated still more, the viscous mass becomes thin and pours readily, and at about 450° C. it boils. The yellowish vapor arising may be condensed, forming "flowers of sulphur."

These different liquids also might be called allotropic modifications of sulphur; in fact, the solid and molten form of every element might be called allotropic modifications of that element, but it is not customary to use the term for forms of substances which change into one another as readily as these do.

Sulphur is used in vast quantities in making gunpowder, matches and fireworks, in vulcanizing rubber, and in the manufacture of sulphuric acid and other sulphur compounds.

COMPOUNDS OF SULPHUR.

Sulphur Dioxide.

When sulphur is burned in oxygen a gaseous compound is formed, called sulphur dioxide.* Determinations of its composition and density show its formula to be SO_2. It dissolves in water, giving an acid as shown by experiment 4, p. 31. This is called sulphurous acid, and the formula H_2SO_3, $(SO_2 + H_2O)$ is assigned to it to correspond with a series of salts called sulphites which may be formed from it. Examples are sodium sulphite, Na_2SO_3, and potassium sulphite, K_2SO_3. There are other salts known called *bisulphites* or *acid sulphites*, such as potassium bisulphite, $KHSO_3$. This illustrates a very common property of dibasic acids, namely, their ability to form two kinds of salts with the same metal, in which either one or both reacting weights of hydrogen are replaced by the metal. In the case of the other dibasic acid that has been described, namely, carbonic acid, there are also two series of salts, carbonates and bicarbonates. Sodium carbonate, Na_2CO_3 (a hydrate of which, $Na_2CO_3\ 10\,H_2O$, is the common washing soda), and sodium bicarbonate or baking soda, $NaHCO_3$, are important examples.

Preparation of Sulphur Dioxide. Although, commercially, sulphur dioxide is prepared by burning sulphur, or certain of its compounds, a convenient laboratory process may be based on the decomposition of any sulphite by sulphuric acid.

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* In accordance with a common usage, sulphur dioxide is sometimes called sulphurous anhydride, the term *anhydride* being given to an oxide which with water gives an acid. Another example of this use of the term is nitric anhydride for nitrogen pentoxide, etc.

SUPPLIES: Sodium sulphite, sulphuric acid, splinter, colored flower and piece of colored calico.

Experiments:

1. Put some sodium sulphite in the bottom of a jar or bottle and add slowly dilute sulphuric acid. Smell the gas that is evolved.

2. Hold a burning splinter in the bottle containing the gas.

3. Prepare a fresh bottle full of the gas, and when effervescence has ceased, suspend a colored flower, *e.g.*, a head of red clover, or a piece of *moist* colored calico in the bottle and cover it with a glass plate.

4. Pour some of the gas, prepared in the above manner, into an empty jar, being careful not to pour any of the sulphuric acid with it, add a little water, place the hand over the mouth of the jar and shake. Leave the solution in the jar exposed to the atmosphere for several days, until the chapter on sulphuric acid has been read.

Notes on Sulphur Dioxide.

Sulphur dioxide is a colorless gas having a suffocating odor. It does not burn or support combustion, and is about 2.2 times heavier than air. One volume of water at ordinary temperature will dissolve about fifty volumes of the gas.

Sulphur dioxide is extensively used as a *bleaching agent.* It bleaches by deoxidation (reduction), and therefore bleaches those coloring matters rich in oxygen which part with all or a portion of their oxygen. Silk, hair, straw, wool and many other substances, which would be injured by chlorine, are whitened by sulphur dioxide. Objects bleached by sulphur dioxide usually become yellow with age, as the substances formed by reduction slowly become oxidized again by the action of the oxygen in the air.

The gas is also used as a *disinfectant*, as it destroys vegetable organisms. Its destructive effect on plant life may be noticed in the neighborhood of smelters and chemical works in which sulphur dioxide is generated, and for miles around our nickel mines, where the ore containing sulphur is piled in heaps and roasted. It is used to fumigate houses and clothing, in the "curing" of meats, wines and beer, for purpose of destroying mould and other organisms, in the manufacture of paper, in tanning, in refining sugar, and in the manufacture of sulphuric acid (see pp. 197 and 201). Liquid sulphur dioxide is now on the market in metal cylinders.

The bleaching action of sulphur dioxide is a process which takes place only in the presence of moisture. This action of water as a *catalyser* for reactions involving gases, is by no means an uncommon phenomenon. Perfectly dry ammonia and hydrogen chloride, for instance, have no action on one another, while the merest trace of water vapor is sufficient to cause the immediate formation of the white fumes of ammonium chloride seen in experiment 5, p. 182. These gases prepared in the laboratory generally contain sufficient moisture to cause reaction, unless special precautions are taken to dry them.

Solutions of salts of sulphurous acid, both sulphites and bisulphites, are also reducing agents, and on account of this property have an extensive use in photography.

Sulphur Trioxide.

Although sulphur dioxide is a powerful reducing agent, it reacts, under ordinary conditions, only very slowly with the oxygen of the atmosphere. Certain catalysers, however, cause the oxidation of sulphur dioxide by oxygen or air to take place rapidly, a product being formed with a composition represented by SO_3, and consequently called sulphur trioxide. One of the most effective catalysers for this purpose is finely divided platinum, used in the form of "platinized asbestos."

If some platinum asbestos can be obtained* this process may be illustrated as follows :—

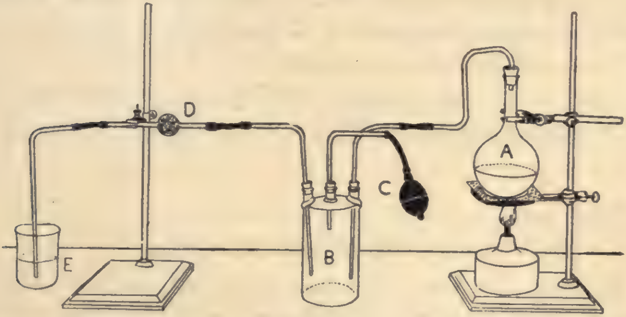

FIG. 44.

Preparation of Sulphur Trioxide. Arrange apparatus as in Fig. 44. A is a flask for generating sulphur dioxide from copper clippings and strong sulphuric acid. The delivery tube leads to a three-necked Woulff bottle B, into which air is being forced by means of an aspirator bulb C. From there the mixed gases are forced over heated platinized asbestos in the hard glass bulb tube D, and pass into the beaker E.

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* It may be prepared by moistening asbestos with a solution of platinum chloride, drying, moistening again with a strong solution of ammonium chloride, drying and heating to redness.

The Woulff bottle serves as a chamber in which the sulphur dioxide and air are mixed, and the platinized asbestos acts as a catalyser, promoting the union between the sulphur dioxide and oxygen of the air. Dense white fumes of sulphur trioxide are seen to fill the beaker; after these have been shown, the fumes should be led into water, and the solution tested for sulphuric acid by barium chloride.

Like the dioxide, sulphur trioxide dissolves in water, forming an acid called sulphuric acid; for this reason it is often called sulphuric anhydride. When first formed it is a colorless, mobile liquid which solidifies at 15° C. and boils at 46° C., but this, in time, changes to a white, fibrous, crystalline mass. When exposed to moist air it forms dense fumes consisting of minute drops of sulphuric acid. Its reaction with water is accompanied with great evolution of heat and a hissing sound.

SULPHURIC ACID, H_2SO_4.

SUPPLIES: Test tube, litmus paper, wood shavings, white sugar, starch, sweet oil, strong sulphuric acid; solutions of lead acetate, sodium carbonate, barium chloride and some sulphate; copper and zinc.

Experiments:

1. Add a few drops of concentrated sulphuric acid to a little water in a test tube. Note the change of temperature.

Nearly fill the tube with water and taste the solution. Test with litmus. To a portion add a piece of zinc. To another portion add a small quantity of sodium carbonate.

2. Place a drop of concentrated sulphuric acid on a wood shaving, a piece of paper, some white sugar, or starch. What element is present in each of these substances?

3. Shake some concentrated sulphuric acid and sweet oil together in a test tube. What element is present in the oil?

4. To a solution of a lead salt add a drop or two of dilute sulphuric acid.

5. Repeat experiment 4, using barium chloride instead of the lead salt.

The precipitate formed in this experiment is insoluble in hydrochloric acid, and its formation is good evidence of the presence of sulphuric acid, or of a sulphate.

6. Test the solution left from experiment 4, p. 195, by a few drops of barium chloride. Does the solution contain sulphuric acid?

7. Heat a little concentrated sulphuric acid with a piece of copper in a test tube. Smell the gas given off. What is it?

8. Repeat 7, using a piece of granulated zinc in place of the copper.

The difference between the action of the concentrated and that of the dilute sulphuric acid on zinc (see p. 56) should be noted.

Experiment 1 illustrates some of the properties which sulphuric acid possesses in common with other acids. It also shows the large amount of heat produced when the solution is diluted.

Experiments 2 and 3 afford examples of a property possessed by concentrated sulphuric acid, of reacting with many substances containing oxygen and hydrogen, by which these elements are removed from the compound in the proportion in which they are present in water, thus forming a more dilute acid and leaving the other constituents of the compound. In the above cases a residue of carbon was found.

The result of experiment 6 shows that the sulphurous acid has been oxidized to sulphuric acid by the air. This change corresponds to the oxidation of sulphur dioxide to sulphur trioxide. The process shown in experiments 7 and 8, in which the sulphuric acid was reduced by the metal to sulphur dioxide, is the reverse of this. It is analogous to the reduction of nitric acid by copper, Exp. 2, p. 176.

This reaction between copper and concentrated sulphuric acid gives a convenient laboratory method for preparing sulphur dioxide, and will be made use of in the experiment described below.

The rate of evolution of the gas may be controlled by regulating the temperature of the flask, the rate of reaction increasing with rise of temperature, in accordance with a general principle, numerous illustrations of which have already been met with.

In the older commercial process for the manufacture of sulphuric acid, the oxidation of the sulphur dioxide by the oxygen of the air is effected by the use of oxides of nitrogen, which hasten the combination of the oxygen and the sulphur dioxide, much in the same way as the platinized asbestos did in the experiment described above. The sulphur dioxide is generally produced by burning a compound of iron and sulphur, known as iron pyrites. The gas is led into large leaden chambers along with air and the oxides of nitrogen. Steam is also admitted to the chamber, and reacts with the sulphur trioxide, forming sulphuric acid which collects at the bottom of the chamber, and is afterwards concentrated by evaporation.*

This process may be illustrated very imperfectly in the laboratory as follows :—

Experiment :

Arrange apparatus as shown in Fig. 45.

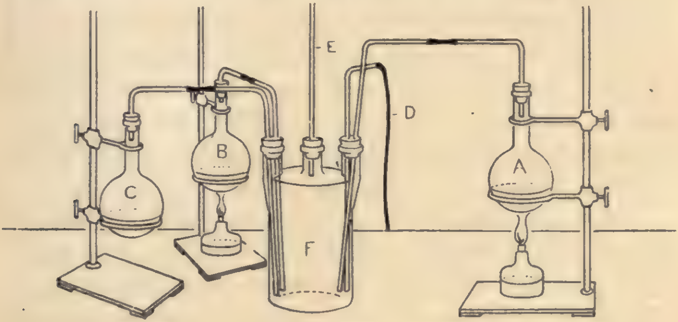

Fig. 45.

A is a flask from which steam is generated.

B is a flask containing copper clippings and sulphuric acid for generating sulphur dioxide.

C is a flask containing copper and nitric acid for generating nitric oxide.

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* Although the manufacture of sulphuric acid is of no greater importance than many other commercial applications of chemistry, the details of this or some other typical process should be given the pupils, in order to show the difference between a laboratory experiment and the technical application of the same reaction. Their attention should be especially drawn to the importance of producing the sulphur dioxide from the cheapest sources. In some places, for instance, iron pyrites is the cheapest source of this gas, while in others sulphur or sulphur compounds may be at hand as waste products from some other manufacture. The ingenuity exercised in the recovery of the oxides of nitrogen should be pointed out and the necessity of working up, as far as possible, all by-products should be emphasized. For all such details the teacher must refer to some standard treatise on chemistry.

D is a tube through which air is blown into the three-necked Woulff bottle, F.

E is an escape tube for waste gases.

Pass the steam, nitric oxide, sulphur dioxide and air into the Woulff bottle, which acts as a condensing chamber. Sulphuric acid collects at the bottom of the bottle.

The reactions are complex and variable. The following explanation of the action of the oxides of nitrogen has been given:—

The nitric oxide may be supposed to act as a "carrier" of oxygen. It combines with the oxygen of the air, forming nitrogen peroxide, NO_2, which gas acts as an oxidizing agent, converting the sulphurous acid into sulphuric acid, or the sulphur dioxide into sulphur trioxide.

Since the nitric oxide is found at the beginning and end of the process in the same condition, and is not used up, the action may be called a *catalytic* one.

It should be noted that this experiment does *not* illustrate the manufacture of sulphuric acid, but merely the reaction between sulphur dioxide, air, nitric oxide, and steam. In the first place, we have used up in the preparation of the sulphur dioxide the very material that we were making--namely, sulphuric acid. Secondly, the oxides of nitrogen have been produced by a method that would cost more than the value of sulphuric acid obtained, and finally, no attempt was made to recover the waste products.

Another commercial method for the manufacture of sulphuric acid which was introduced a few years ago and is rapidly gaining ground, is based on the catalytic action of platinum mentioned above (see p. 197). It is known as "the contact process."

Heat of Reaction.

It will have been noticed that reactions are almost invariably accompanied by the production or using up of heat, which when the reaction takes place quickly is indicated by a change of temperature. The large amount of heat produced when water was added to concentrated sulphuric acid was shown in Exp. 1, p. 198. The following modification of this experiment will illustrate a very important general law which has been found to govern the amount of heat used up or produced in reactions.

Experiment:

Into a beaker measure out accurately 20 cc. of strong ammonia solution, and add 50 cc. of water. Into another beaker measure out 10 cc. of concentrated sulphuric acid; to this add 50 cc. water, pouring it in quickly, to prevent the acid spurting, and note the temperature immediately with a thermometer. Without further delay pour this liquid into the other beaker containing the ammonia; stir, and again note the temperature.

We have thus prepared a solution of ammonium sulphate in two stages. In the first place, sulphuric acid was diluted, and in the second place, this diluted solution was allowed to react with the ammonia solution. Each of these stages was accompanied by the production of heat, the total amount of heat being proportional to the total rise in temperature multiplied by the number of grams of liquid heated.

Now we can arrive at the same final result in a different way. Measure out again 20 cc. of strong ammonia, but this time add to it 100 cc. of water. To this add carefully 10 cc. of concentrated sulphuric acid, pouring it slowly down the stem of the thermometer with which the liquid is being stirred. Note the temperature to which the liquid rises, and compare with the final temperature of the first experiment.

It will be noticed that in both experiments the same volume of the same product was obtained, but that in the second experiment the intermediate stage of diluting the sulphuric acid was omitted. The final temperature was the same in both cases, which shows that the same total quantity of heat was produced, irrespective of the method employed for obtaining the final product from the original substance.

This illustrates a general law. The total amount of heat produced or used up in any reaction depends only on the original substances and the final product, and is independent of the method employed for effecting the change. If the process can be reversed, the same amount of heat that was produced in the original change will now be used up, and *vice versa*, if heat has been used up in the first case, the same amount will be produced in the reverse process.

In speaking of slow combustion (p. 40), it has already been pointed out that the total evolution of heat is independent of the rate of the reaction. The total heat produced by the combustion of a ton of coal is the same whether the fire burns quickly or slowly, although, when the fire burns brightly, higher temperatures are obtained for a shorter time.

In stating this law it is assumed that nothing else is produced during the process that may take the place of heat, or if there is, that the corresponding amount of heat must be allowed for. As is well known, heat may be converted into mechanical work, as in the steam engine, or into electrical energy, as in the power house. Consequently if mechanical work, or electricity, is produced during a reaction, the amount of heat will be correspondingly less.

Notes on Sulphuric Acid.

Sulphuric acid is an oily liquid, colorless when pure, but sometimes darkened by organic matter. Frequently the bottom of bottles containing sulphuric acid is covered with a white precipitate of lead sulphate, due to the slight action of the acid on the lead lined chamber in which it is manufactured. When sulphuric acid is added to water, a reaction takes place and heat is produced. The concentrated acid abstracts the elements of water from any organic matter, such as wood, paper, cloth, sugar, starch, flesh, etc., leaving the carbon. Concentrated sulphuric acid is also used for "drying" gases which contain water vapor.

Sulphuric acid is extensively used in the manufacture of other acids, as nitric, hydrochloric, etc. It is used in the manufacture of washing soda, Na_2CO_3 $10\,H_2O$, baking soda, $NaHCO_3$, alum, glucose, nitroglycerine, dyestuffs and fertilizers; it is also used in processes of dyeing, bleaching, electroplating, and in refining oils.

Sulphates.

Sulphuric acid forms numerous and important salts. Among these may be mentioned sodium sulphate, Na_2SO_4, potassium sulphate, K_2SO_4, ammonium sulphate, $(NH_4)_2SO_4$, calcium sulphate, $CaSO_4$, barium sulphate, $BaSO_4$, and copper sulphate, $CuSO_4$. Many of these salts crystallize from their aqueous solutions with "water of crystallization" (see p. 75), giving salts with the following formulae:—Glauber's salt, $Na_2SO_4\ 10\,H_2O$, bluestone, $CuSO_4\ 5\,H_2O$, etc. These hydrated salts are chemical individuals, as stated on page 84, because the crystals obtained from a saturated solution at the ordinary temperature are identical in composition with those obtained from the saturated solution at a higher or a lower temperature, although the concentrations of these solutions may be very different. Often, however, if the temperature is altered sufficiently, crystals of a different composition are obtained, but these again have constant composition over a wide range of various solutions and are also chemical individuals, though different from the first. (Compare the two compounds of carbon and oxygen, carbon monoxide and carbon dioxide.)

As is characteristic of all chemical individuals, the proportions of the constituents are simple multiples of the reacting weights. For example, bluestone contains 159.66 g. of anhydrous copper sulphate to 90.1 g. of water, or one reacting weight of copper sulphate to five reacting weights of water. If a few grams of bluestone be powdered finely, weighed, and heated for some time on a watch glass to a temperature of about 110° C. (or less, depending on the amount of moisture in the atmosphere), it will be found to have lost weight, owing to water vapor having been given off. The substance remaining has the composition represented by the formula $CuSO_4H_2O$, four reacting weights of water having been given off by one reacting weight of bluestone. The substance, $CuSO_4H_2O$, or copper sulphate monohydrate, on heating to a still higher temperature (about 200°), decreases further in weight, the loss corresponding to the change of the substance to one with the composition $CuSO_4$, or "anhydrous copper sulphate." This case is mentioned merely because it is typical of the behavior of hydrated salts, and furnishes a good illustration of the law of reacting weights.

There is also a series of bisulphates analogous to the bisulphites and bicarbonates (see p. 194), such as potassium bisulphate $KHSO_4$.

Test for Sulphuric Acid and Solutions of Sulphates: To solutions of sodium sulphate, ammonium sulphate, potassium sulphate and copper sulphate, add a few drops of a solution of barium chloride and a little hydrochloric acid.*

Compare the precipitate with that formed in Exp. 5 with sulphuric acid. The substance formed is an insoluble salt of sulphuric acid, namely, barium sulphate. Its formation serves as an indication that a solution contains either sulphuric acid or a sulphate.

Hydrogen Sulphide, H_2S.†

SUPPLIES: Apparatus as in Fig. 33, ferrous sulphide, hydrochloric or sulphuric acid, litmus paper, carbon disulphide, glass tube drawn to a fine point.

Preparation. Arrange apparatus as shown in Fig. 33. Place in the test tube some ferrous sulphide, cover it with water and add a little hydrochloric or sulphuric acid. Collect a bottle of the gas by displacement of air, and then dip the delivery tube in a bottle filled with water, so as to prepare a solution of the gas.

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* Solutions of many other substances, such as sodium carbonate, give a precipitate with barium chloride solution, but not when hydrochloric acid is present.

† Hydrogen sulphide has a very offensive odor and is poisonous when breathed in quantity, therefore the preparation of this gas and all experiments with it should be performed in a fume cupboard.

Experiments with Hydrogen Sulphide:

1. (*a*) Cautiously smell the gas collected in the bottle; (*b*) test it with moistened litmus paper; (*c*) bring a lighted match to the mouth of the bottle

What collects on the sides of the bottle? Carefully smell the gas that is left in the bottle. What is it?

2. Repeat the preparation of hydrogen sulphide, but replace the delivery tube by a glass tube drawn to a point. After assuring yourself that all the air is out of the apparatus (see Hydrogen, p. 56), light the gas as it issues from the tube and bring into the flame a cold object, such as an evaporating dish.

What substances are deposited on the cold surface? What elements are shown to be present in the gas by these experiments?

Hydrogen sulphide is a compound of hydrogen and sulphur, represented by the formula H_2S. It burns in the air, giving sulphur dioxide and water vapor, but when the supply of air is limited sulphur is produced. Sulphur is also formed if the temperature of the combustion is kept low, as when a cold object is inserted in the flame.

The gas dissolves in water, forming a solution with weak acid properties and marked reducing power.

Hydrogen Sulphide Water, or Hydrosulphuric Acid.

SUPPLIES: Litmus, magnesium, zinc, solution of potassium permanganate, nitric acid; solutions of any salt of lead, silver, copper, iron (ferrous), arsenic, antimony, zinc and magnesium; silver or copper coin.

Experiments:

With the solution of hydrogen sulphide, make the following experiments;

1. Test with litmus paper.

2. Add a freshly cleaned strip of magnesium.

3. Add a piece of zinc.

Is any hydrogen evolved? Are the acid properties strongly marked?

4. Add a little very dilute solution of potassium permanganate. (The disappearance of color indicates that the latter substance has been reduced.)

5. Add a little strong nitric acid and warm.

What do the brown fumes indicate? What is the precipitate?

On account of its acid properties the solution is sometimes called hydrosulphuric acid. It forms a number of salts, called sulphides, most of which are only very slightly soluble in water, and many of which have characteristic colors which are useful in separating and identifying the metals.

6. To a solution of any salt of (*a*) lead, (*b*) silver, (*c*) copper, (*d*) iron (ferrous), (*e*) arsenic, (*f*) antimony, (*g*) zinc, (*h*) magnesium, add some hydrosulphuric acid, or bubble hydrogen sulphide gas through the solution.

In which cases are precipitates formed?

7. Place a drop of hydrosulphuric acid on a clean silver or copper coin. Compare the action on magnesium in experiment 2.

The formulae of many of the sulphides, PbS, CuS, etc., may be remembered by noting that hydrosulphuric acid is a dibasic acid, with the formula H_2S. There are, however, other compounds of sulphur with metals which cannot be derived from this formula.

Some sulphides, such as ferrous sulphide, used in the preparation of the gas, react with acids, such as hydrochloric or sulphuric, giving off hydrogen sulphide and forming the chloride or sulphate of the metal, as represented by the equations

$$FeS + 2\,HCl = FeCl_2 + H_2S,$$
$$FeS + H_2SO_4 = FeSO_4 + H_2S.$$

Notes on Hydrogen Sulphide.

Hydrogen sulphide, also known as sulphuretted hydrogen, is a colorless gas having a very unpleasant odor. The odor of rotten eggs is due to the presence of this gas, which is formed during the putrefaction of organic matter containing sulphur. For this reason hydrogen sulphide is also present in gases given off from sewers and cesspools. It occurs in volcanic gases and in the water of some mineral springs. As seen from experiment 7, hydrosulphuric acid (or, what amounts to the same thing, hydrogen sulphide in the presence of moisture), rapidly tarnishes certain metals, owing to the formation of sulphides. For this reason, brass apparatus should not be exposed to the action of this gas in the laboratory.

CHAPTER XVI.—STRENGTH OF ACIDS AND BASES

SOME acids, even in concentrated solution, do not possess the characteristic acid properties to the same extent as hydrochloric, nitric, sulphuric acids, etc., which are called "strong" acids. Carbonic acid, for instance, reacts but slowly with a piece of magnesium; its taste is very faintly sour, and its action on litmus is not so marked as is that of the three strong acids mentioned. Acids, such as carbonic, sulphurous, and hydrosulphuric, whose acid properties are thus weakly marked, are called "weak" acids. A similar distinction is made between strong and weak bases, sodium, potassium and calcium hydroxides being examples of strong bases.* In both cases the difference is only one of degree, however, and acids and bases of intermediate strengths are known.

An important respect in which the difference in "strength" is shown, is in the properties of solutions of salts. A salt like sodium carbonate, formed from a weak acid, carbonic acid and a strong base, sodium hydroxide, when dissolved in water, gives a solution with *basic* properties, while the solution of a salt corresponding to a weak base and a strong acid shows *acid* properties.

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* The hydroxides of copper, iron, and aluminium are spoken of as bases, although they are almost insoluble in water.

Experiments :

Test solutions of the following salts with litmus paper: Sodium carbonate, potassium carbonate, sodium sulphite, sodium acetate, copper sulphate, ferrous sulphate, and alum.

Point out the weak acid, or weak base, corresponding to each salt.

That hydrogen chloride, nitric acid and hydrogen sulphide can be prepared by treating chlorides, nitrates and sulphides with concentrated sulphuric acid, must not be taken to mean that sulphuric acid is *stronger* than hydrochloric, nitric or hydrosulphuric acid, in the sense in which the word is used above. The fact that the substances in question may be prepared in this way is due chiefly to their volatility.

The Properties of Dilute Aqueous Solutions.

It has been pointed out that gaseous solutions differed from other solutions in that their properties might be calculated from the properties of the constituents by a very simple rule.

Dilute aqueous solutions of certain substances (strong acids and bases and salts) also differ from all other solutions in a somewhat similar way. We shall illustrate this by reference to the acids and bases already met with.

The dilute solutions of hydrochloric, sulphuric and nitric acids have a set of common properties repeatedly referred to as "acid properties." Besides this, however, each acid has other properties which are different in the different acids, but which are common to all the salts of that acid. For example, hydrochloric acid will give a precipitate of silver chloride with a solution of silver nitrate, and this property, as shown on p. 147, is common to all soluble chlorides of the metals.

We express these facts by saying that the solutions of all acids contain the hypothetical substance "hydrogenion," with the constant properties of sourness, colorlessness, ability to turn blue litmus red, etc., and another substance with the properties characteristic of the particular acid in question, *e.g.*, in the case of hydrochloric acid "chloridion," which is colorless, has the property of precipitating silver nitrate solution, etc., and is common to all solutions of chlorides. By the assumption of *two* such constituents for each substance we can express the characteristic properties of dilute aqueous solutions of it.

The word "hydrogenion," representing the hydrogen of the acid, is better than "hydrogen," since it does not imply that the properties expressed are those of hydrogen gas, whose solutions are not scur and have no effect on litmus. Similarly, the word "chloridion" is employed to prevent confusion with chlorine gas, whose solutions are yellowish and indeed have none of the properties of chlorides.

In the same way the properties of the bases may be conveniently expressed by saying that they contain "hydroxidion," and another "ion," as these imaginary substances are called. For example, sodium hydroxide solution is said to contain sodiumion and hydroxidion, the latter expressing all the characteristic basic properties, and the former those common to sodium salts.

It is thus seen, that by knowing the properties of a comparatively small number of "ions," the properties of the aqueous solutions of a great many acids, bases and salts may be easily remembered.

This is illustrated by experiment 6 with hydrosulphuric acid. It was there stated that *any* salt of lead, etc., might be used. The reason for this is, that all salts of lead behave in the same way with hydrosulphuric acid, a fact which is expressed by saying that they all contain "leadion."

TABLE OF REACTING WEIGHTS OF ELEMENTS.

Aluminium	Al	27.1
Antimony	Sb	120.2
Argon	A	39.9
Arsenic	As	75.0
Barium	Ba	137.4
Bismuth	Bi	208.5
Boron	B	11
Bromine	Br	79.96
Cadmium	Cd	112.4
Caesium	Cs	132.9
Calcium	Ca	40.1
Carbon	C	12.00
Cerium	Ce	140.25
Chlorine	Cl	35.45
Chromium	Cr	52.1
Cobalt	Co	59.0
Columbium	Cb	94
Copper	Cu	63.6
Erbium	Er	166
Fluorine	F	19
Gadolinium	Gd	156
Gallium	Ga	70
Germanium	Ge	72.5
Glucinum (Be)	Gl	9.1
Gold	Au	197.2
Helium	He	4
Hydrogen	H	1.01
Indium	In	115
Iodine	I	126.97
Iridium	Ir	193.0
Iron	Fe	55.9
Krypton	Kr	81.8
Lanthanum	La	138.9
Lead	Pb	206.9
Lithium	Li	7.03
Magnesium	Mg	24.36
Manganese	Mn	55.0
Mercury	Hg	200.0
Molybdenum	Mo	96.0
Neodymium	Nd	143.6
Neon	Ne	20
Nickel	Ni	58.7
Nitrogen	N	14.04
Osmium	Os	191
Oxygen	O	16.00
Palladium	Pd	106.5
Phosphorus	P	31.0
Platinum	Pt	194.8
Potassium	K	39.15
Praseodymium	Pr	140.5
Radium	Ra	225
Rhodium	Rh	103.0
Rubidium	Rb	85.5
Ruthenium	Ru	101.7
Samarium	Sm	150.3
Scandium	Sc	44.1
Selenium	Se	79.2
Silicon	Si	28.4
Silver	Ag	107.93
Sodium	Na	23.05
Strontium	Sr	87.6
Sulphur	S	32.06
Tantalum	Ta	183
Tellurium	Te	127.6
Terbium	Tb	160
Thallium	Tl	204.1
Thorium	Th	232.5
Thulium	Tu	171
Tin	Sn	119.0
Titanium	Ti	48.1
Tungsten	W	184
Uranium	U	238.5
Vanadium	V	51.2
Xenon	Xe	128
Ytterbium	Yb	173.0
Yttrium	Yt	89.0
Zinc	Zn	65.4
Zirconium	Zr	90.6

SUITABLE CONCENTRATIONS FOR LABORATORY SOLUTIONS.

	Formula	*Grams per litre*
Dil. hydrochloric acid	HCl	150
Dil. sulphuric acid	H_2SO_4	300
Dil. nitric acid	HNO_3	300
Potassium bichromate	$K_2Cr_2O_7$	80
Ferric chloride	$FeCl_3$	10
Ferrous sulphate	$FeSO_4\ 7\ H_2O$	50
Potassium iodide	KI	15
Calcium hydrate	$Ca(OH)_2$....Saturated solution.	
Calcium chloride	$CaCl_2$	75
Calcium sulphate	$CaSO_4\ 2\ H_2O$.....Saturated solution.	
Magnesium chloride	$MgCl_2\ 6\ H_2O$	30
Ammonium oxalate	$(NH_4)_2C_2O_4\ H_2O$	30
Sodium phosphate	$Na_2HPO_4\ 12\ H_2O$	60
Mercuric chloride	$HgCl_2$	60
Lead acetate	$Pb(C_2H_3O_2)_2\ 3\ H_2O$	80
Silver nitrate	$AgNO_3$	15
Ammonium chloride	NH_4Cl	100
Ammonium carbonate	$(NH_4)_2CO_3$	100
Potassium hydroxide	KOH	100
Sodium hydroxide	$NaOH$	100
Barium chloride	$BaCl_2\ 2\ H_2O$	100
Sodium carbonate	$Na_2CO_3\ 10\ H_2O$	100
Ammonium sulphocyanate	NH_4SCN	100

APPENDIX

THE ATOMIC HYPOTHESIS.

When new and striking phenomena are first discovered, we often try to connect them together in the form of a mental picture by assuming some imaginary cause for the existence of the facts. For example, a ruddy glow above the horizon may at once suggest the mental picture of a burning house, and this assumption would afford a possible explanation of the light in the sky. Having made such an assumption, we are apt to think that we have "explained" the facts. We may draw conclusions as to the location of the building from the position of the light, and as to its size from the intensity of the glow. Absence of steam and smoke may be even taken as "proof" that the fire brigade has not yet arrived on the scene. On hearing next day that the same phenomenon had been observed all over the country, we give up our mental picture of the burning house and content ourselves by saying that we had seen a rather unusual type of the Aurora Borealis.

Such mental pictures have, from time to time, been made in science. One of these, the atomic hypothesis, has had such a great influence on the past history and present nomenclature of chemistry that a short account of it may not be out of place. In order to "explain" the facts expressed in this book by the "law of reacting weights," Dalton, in 1803, made the assumption that matter consisted of small indivisible particles—the atoms—which in compound substances were combined together in small groups to form compound particles. By his further assumption that the atom of each substance had a small but constant weight, proportional to the combining weight, the facts of the law of reacting weights were thus "explained." According to this hypothesis carbon dioxide, for example, was assumed to consist of atoms, each containing one atom of carbon and two atoms of oxygen, while the atoms of carbon monoxide were supposed to contain one atom of carbon and one atom of oxygen each. Soon after this, chemical symbols were introduced; the symbol C, for example, stood for an atom of carbon, the symbol O for an atom of oxygen, etc. Thus, a symbol that is now employed to represent a definite number of grams of a particular substance, without reference to any hypothesis, was originally introduced to indicate the hypothetical atom of that substance.

From time to time, as new facts were discovered, attempts were made to "explain" these also by modifying and extending the atomic hypothesis. The relation between the volumes and reacting weights of gases led to the introduction of the hypothetical "molecule" and Avogadro's hypothesis; and the pressure and temperature laws of gases were accounted for by the idea that these molecules were flying about with great rapidity. Again, just as the assumption of the burning house necessitated further assumptions as to its locality and size, so also the development of the molecular hypothesis led to a determination of the size of the molecules.

Later still, certain phenomena observed in connection with aqueous solutions necessitated an even more complicated modification, namely the assumption that some molecules break up into parts endowed with electrical charges and called ions; while quite recently, owing to new discoveries in radioactivity which do not appear to fit in with the idea of the indivisibility of the atom, it has been assumed that the atoms themselves are capable of subdivision into a large number of smaller particles, to which the name electrons has been given.

At the present day, now that the laws governing the fundamental facts are found to be so simple and general, there is a tendency among chemists to abandon these cumbrous hypotheses, not only on account of the greater simplicity of the facts themselves, but also because it is now recognized that such hypotheses, although of great service in the past, have often led to unclear expression of the facts and one-sided development of the science. Just as, for instance, the burning house hypothesis may have prevented some people from studying a rare specimen of the Aurora, so also the atomic hypothesis is responsible for the imperfect study or total neglect of certain classes of phenomena which did not fit in well with this explanation.

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